

**practising
science**



SXR103

Practising Science: study book



SXR103 Course Team

<i>Academic contributors:</i>	Evelyn Brown (chair) Roger Beck Mary Bell Steve Blake Bob Hill Mark Jones Sally Jordan Jean McCloughry Pat Murphy Valda Stevens Peter Taylor Ruth Williams
<i>Course Manager:</i>	Elaine McPherson
<i>Course Secretary:</i>	Elsie Frost
<i>Course Clerk:</i>	Donna Gibb
<i>Project Officer:</i>	Pete Wood
<i>Editor:</i>	Dick Sharp
<i>Indexer:</i>	Jane Henley
<i>Graphic Artist:</i>	John Taylor
<i>Graphic Design:</i>	Glen Darby, Jenny Nockles
<i>Library Liaison:</i>	Lydia Eaton
<i>External Assessor:</i>	David Hill

This publication forms part of an Open University course SXR103 *Practising Science*. Details of this and other Open University courses can be obtained from the Course Information and Advice Centre, PO Box 724, The Open University, Milton Keynes MK7 6ZS, United Kingdom: tel. +44 (0)1908 653231, e-mail ces-gen@open.ac.uk

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First published 2001; second edition 2003.

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Edited, designed and typeset by The Open University.

Printed and bound in the United Kingdom by The Alden Group, Oxford.

ISBN 0 7492 5852 7

SXR103 study book i2.1.

Contents

1	Introduction	5
1.1	Science and scientists	5
1.2	Introducing the <i>Practising Science study book</i>	5
1.3	What to do next	6
2	Earth Sciences: reading the rocks	7
2.1	Introduction	7
2.2	Minerals and rocks	8
2.3	The formation of igneous rocks	10
2.4	The formation of sedimentary rocks	16
2.5	The formation of metamorphic rocks	24
2.6	Interlude	27
2.7	Geological fieldwork	27
2.8	The rock cycle	32
2.9	Changing sea-level	33
2.10	What to do next	34
2.11	Summary of Section 2	34
3	The physics of atoms, energy and radiation	37
3.1	Introduction	37
3.2	The building blocks of the world	37
3.3	Spectral lines: an atom's signature	38
3.4	Into the nucleus	47
3.5	Radioactive decay	48
3.6	Self-assessment questions	60
3.7	What to do next	62
3.8	Summary of Section 3	63
4	Biology: the science of life	65
4.1	Activity D: 'The biological effects of gamma radiation'	65
4.2	What is ecology?	68
4.3	Summary of Section 4	78
5	Chemistry: analysis, reactions and structures	80
5.1	Introduction	80
5.2	The reactivity of metals with hydrochloric acid solution	80
5.3	Displacement reactions	84
5.4	Enthalpy change for a displacement reaction	87
5.5	Activity B: 'Analysing our environment'	88
5.6	Summary of Section 5	92

6	Research project: 'Aluminium: exploring its impact'	93
6.1	The purpose of this project	93
6.2	The scope of the project	94
6.3	What to do next	100
	Answers and comments for in-text activities	101
	Answers to self-assessment questions	108
	Appendix 1 Sines, cosines and tangents	112
	Appendix 2 The equation of a straight line	114
	Appendix 3 The Periodic Table	116
	Acknowledgements	118
	Index	119

An introduction to *Practising Science*

1

1.1 Science and scientists

Science is all about knowledge, what we know about the material world and the Universe in which our world is just a microscopic speck. The aim of scientists is to extend the frontiers of this knowledge so that we can understand more about the physical Universe and the life within it.

Scientists acquire knowledge by engaging in four fundamentally important and connected tasks. The first is *observation*: they observe the natural world and the space beyond it, and both describe and record what they see. Second, they *construct hypotheses* to explain what they see. Third, they carry out experiments where possible to test their hypotheses. Finally they *communicate their findings* – to other scientists who will build on this work to extend knowledge still further, to technologists who will devise practical applications for scientific knowledge, and to the general public to raise awareness of scientific discovery. The way in which science is communicated to interested parties is especially important because scientific knowledge is useless if no one else can understand it.

You may well find that practising science in a laboratory or field setting is what you will enjoy most about this course. However, graduate scientists leaving universities must be able to demonstrate that they have gained the skills needed to engage effectively with *each* of the tasks described above. You will be introduced to these skills during the Residential School component of *Practising Science* and you will engage with all of the tasks of the scientist as you work through the Residential School activities.

1.2 Introducing the *Practising Science study book*

You will also need an understanding of some of the basic concepts of science and practical techniques if you want to obtain the maximum benefit from the activities at the Residential School. The aim of the materials in this book is to provide this background. As well as studying this material, it is important that you *read through each of the activity workbooks before you attend the Residential School*, to become familiar with the work you will be asked to carry out. We don't expect you to spend a lot of time on this, or to understand everything in the workbooks. Some of it will make sense only as you carry out an activity. However, familiarity beforehand will allow you to progress more easily through the activities and to appreciate their outcomes more fully. You will find it helpful to re-read the summaries in the Course Guide to gain an overview of each activity before reading the associated workbook.

The sections that follow introduce some of the concepts and practical techniques in Earth sciences, physics, biology and chemistry that underpin the various Residential School activities. They also provide you with practice in extracting information from a scientific article, a skill that will be developed as you learn how to communicate science to others. Remember that there is a separate Glossary containing definitions of the scientific terms introduced in this book. These will be indicated in the sections that follow by a superscript G, e.g. atom^G.

1.3 What to do next

The language of science depends heavily on the use of numbers and symbols. Before moving onto Section 2 of this book, you should look at Chapter 5 in *SGSG*, 'Working with numbers and symbols', to check that you feel confident that you can make appropriate use of symbols, such as chemical symbols and units of measurement.

Earth Sciences: reading the rocks

2

2.1 Introduction

The Earth Sciences concern every part of our planet — from the centre of the metallic core, 6400 km below our feet, through its solid and liquid regions (Figure 2.1) to the outer reaches of the atmosphere. But at this point in SXR103, we will concentrate on the outermost, rocky part of the solid Earth, in other words the rocks of the Earth's crust that form the outermost layer of the lithospheric plates. On the scale of a human lifetime, these rocks and the landscapes of which they are part can seem static and immutable, at least in a country such as Britain where there are no active volcanoes, and earthquakes are infrequent and fairly small. None the less, coastal erosion and the biological degradation of rocks to form soils are reminders that geological processes are occurring 'in our own backyard' and that the Earth's surface is continuously being reshaped.

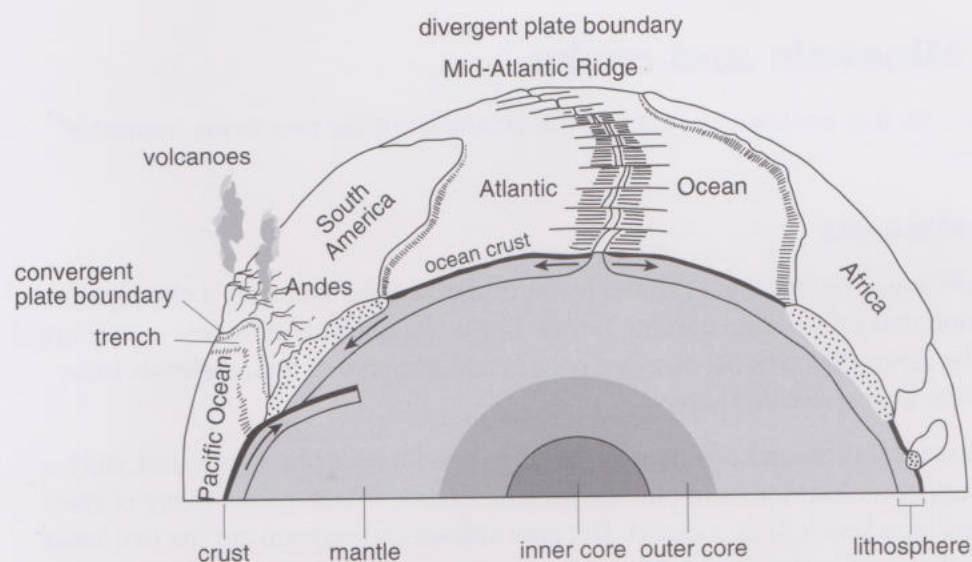


Figure 2.1 A schematic cross-section through the Earth, showing its concentrically layered internal structure and the interacting lithospheric plates on the surface.

Over hundreds of millions of years, the surface of the Earth is rearranged by the slow grind of geological processes. In particular, the rigid outer layer of the Earth (some 100 km thick and known as the lithosphere) is divided into about a dozen plates, which change shape, size and position as their margins experience growth, destruction or collision (see Figure 2.1). So, although the processes can be very slow, given enough time the results include the generation and destruction of mountain ranges that stretch across continents, and the opening and closing of ocean basins. All these geological processes leave their mark by forming new rocks. And those rocks, whether sandstone formed from sand in a desert, or volcanic lava flows, hold clues to the processes that formed them, and therefore the conditions at the time of their formation. In some cases they contain information on past climatic conditions, and some rocks contain fossils that reveal the history of life and the story of evolution. Rocks can provide a narrative of the Earth's history, but we need to learn how to read their tales.

At Residential School you will gain experience in the practical aspects of studying rocks and piecing together the evidence of how they formed. This will involve you in examining rocks exposed in the field (as part of Activity C, 'Investigating the environment') and rock specimens in the laboratory (as part of Activity A, 'Rocks and radioactivity: energy in the Earth'). Both types of investigation are important in geology — the study of rocks and the information that they hold about the history of Earth through its 4.6 billion year history.

In the practical work you will deduce how particular rocks were formed and the sequence of events in the geological history of a small part of the United Kingdom. As with many other areas of science, you'll first gather data by making careful observations. Then, those observations will be interpreted in a way that best explains them.

The aims of this section are to demonstrate how many of the features shown by rocks are inherited from the processes that formed the rock, to introduce you to the skills of sketching features seen in rock exposures, and to show how those features can be interpreted in a simple sequence of events through geological time.

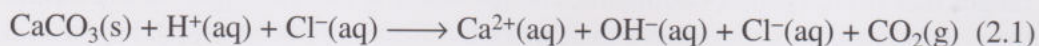
2.2 Minerals and rocks

To begin with, it is necessary to explain the meanings of the two terms 'minerals' and 'rocks'.

2.2.1 Minerals

A mineral^G is a solid material, formed by natural processes and with a chemical composition that falls within certain narrow limits. Its constituent atoms are arranged in a regular three-dimensional array or pattern and because of this, minerals form crystals with characteristic shapes.

Although several thousand different kinds of mineral have been discovered, only a few are very common; for example, the mineral quartz, which forms many of the sand grains on a beach or in a desert. Because silicon and oxygen are the two most abundant elements in the Earth's crust, they are the main ingredients of common minerals, and these minerals are known as silicate minerals^G. They include quartz^G, whose chemical composition is silicon dioxide (formula SiO_2), and a range of others containing additional common chemical elements. A common mineral that is *not* a silicate is calcite^G; it has the chemical composition calcium carbonate (CaCO_3). Whereas most minerals are identified on the basis of physical characteristics (density, hardness, colour, etc.) calcite can also be readily identified with a chemical test. Calcite reacts with dilute hydrochloric acid, liberating bubbles of carbon dioxide (CO_2) gas in a vigorous fizzing froth. You will see this characteristic reaction in the work you do at Residential School. The chemical reaction is written as



In words, this equation can be expressed as: solid calcium carbonate (calcite) plus hydrogen ions in the acid plus chloride ions in the acid gives calcium ions, hydroxide ions and chloride ions dissolved in solution plus carbon dioxide gas.

You'll get more experience with reading and using chemical equations when you study the chemistry material (Section 5).

2.2.2 Rocks

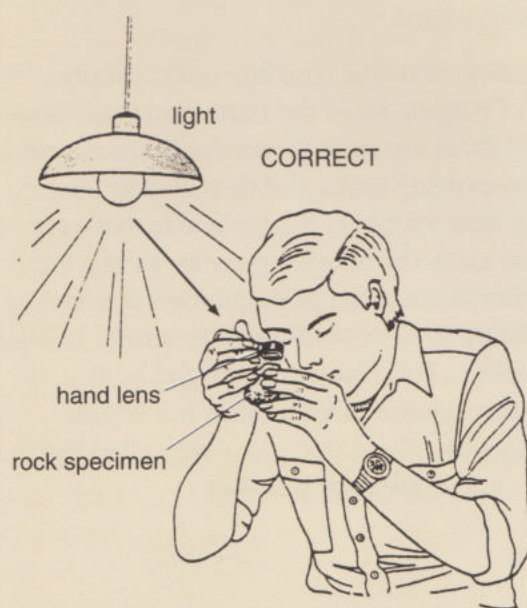
Any naturally formed solid assemblage of mineral grains can be described as a rock^G. The mineral grains may be fragments of crystals or intact crystals and their size can range from a few micrometres (1 micrometre = 10^{-6} m; see *SGSG*, Chapter 5 Section 4.2) to a few centimetres. A rock may consist of one type of mineral but more usually it consists of several minerals. Rocks can be classified according to the way in which the grains are arranged, although the identity of the minerals present (for example, the rock limestone^G is made mostly of calcite), the proportions of particular minerals, and the dominant size of mineral grains are also important. The shape of the grains in a rock, their size and the relationship between them (for example, whether or not the grains interlock with each other to form a mosaic) define the texture^G of a rock, and reflect the processes that formed it. A rock's texture has nothing to do with how the rock feels when you touch it. To 'read the rocks' and discover how any particular rock formed, we investigate its texture and work backwards to deduce the processes by which that texture was produced.

Because the mineral grains in most rocks are quite small it is often best to use a hand lens, typically with a magnification power of times 10, to get a clearer view. You may already have access to a small hand lens. If so, it is a good idea to get some practice in using it before coming to Residential School, and some notes on this are given in Box 2.1, *Using the hand lens*.

Box 2.1 Using the hand lens

The correct way to examine objects with a hand lens is shown in Figure 2.2a. Some additional hints are:

- Hold the hand lens 2–3 cm from your eye and bring the object up towards the lens until it is in focus.
- Make sure the surface of the object is well lit from the side.
- Keep the hand lens and the object parallel to each other and hold both steady.
- With rough surfaces you have to move the object back and forth to bring different parts into focus.



(a)

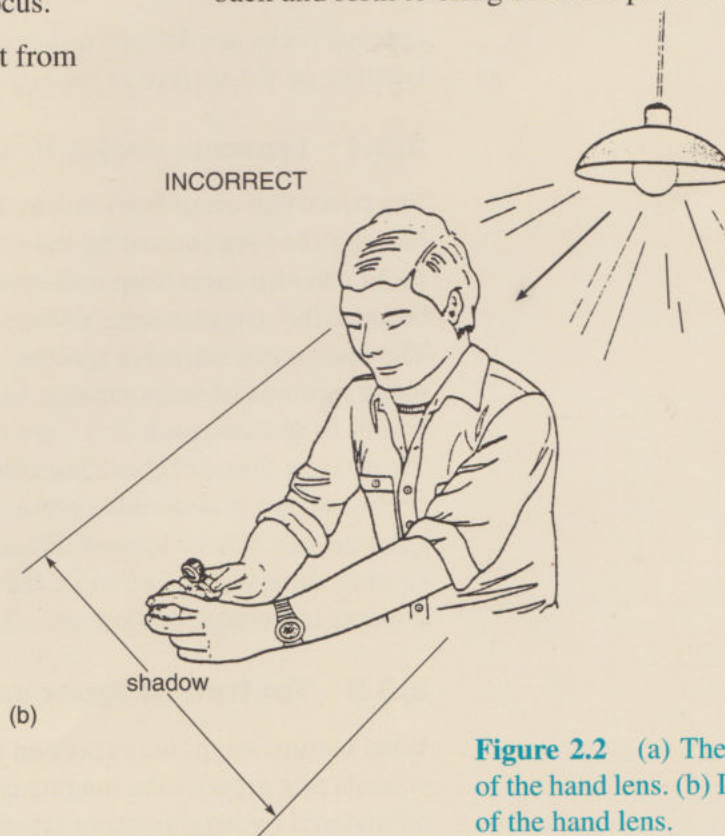


Figure 2.2 (a) The correct use of the hand lens. (b) Incorrect use of the hand lens.

Rocks may be classified into three types, according to the three processes that form rocks:

- 1 Igneous rocks^G: these are formed from molten rock (magma^G) that becomes solid when it cools, either deep underground or after a volcanic eruption at the surface.
- 2 Sedimentary rocks^G: these are formed when eroded particles of pre-existing rocks (in other words sediment, such as sand on a beach or mud on the sea-bed) have been laid down in layers at the surface of the Earth and turned into solid rock by being buried and compacted under more layers of sediment.
- 3 Metamorphic rocks^G: these are existing rocks that have 'changed form' by the action of high pressure or temperature causing new crystals to grow, for example after burial deep in the Earth.

Sections 2.3 to 2.5 describe how each of the three major rock types is formed and how their modes of formation can be deduced from the textures and other features visible in the rocks themselves — reading the rocks. These sections relate to the work you will be doing during the first part of Activity A, 'Rocks and radioactivity: energy in the Earth'. Some of this knowledge will also be used during Activity C, 'Investigating the environment', where you'll be engaged in some fieldwork observing the rocks and interpreting the geological history of an area near your Residential School. After taking stock in Section 2.6, you will be introduced in Section 2.7 to some of the practices of field geology. Section 2.8 is an overview of how the geological processes that form new rocks act on pre-existing rocks in a cycle of destruction, formation and transformation known as the rock cycle^G. Section 2.9 addresses an ongoing geological process, sea-level change, and Section 2.10 refers you to the workbook for Activity C, 'Investigating the environment'. A summary is provided in Section 2.11.

2.3 The formation of igneous rocks

Igneous rocks are defined as having solidified from a molten state, either inside the Earth or on the surface at volcanoes.

2.3.1 Igneous rocks in the landscape

The rocks that erupt from volcanoes are called *extrusive* igneous rocks, simply because they are formed by the extrusion of magma on to the Earth's surface. Igneous rocks can also form deep underground, and these are called *intrusive* igneous rocks, because the magmas were intruded into pre-existing rocks and then slowly cooled. The reason that intrusive igneous rocks are now visible at the surface is that over many millions of years erosion has stripped away the overlying rocks. In this way, bodies of igneous rock that were once subterranean pools of magma are revealed at the surface. Some of these intrusions can be up to several kilometres across. In other cases, magma had intruded pre-existing rocks to form long slab-shaped bodies of igneous rock whose longest dimensions are measured in kilometres but whose shortest dimension is no more than a few metres. These intrusions are called dykes in the case of vertical bodies, and sills in the case of horizontal bodies.

2.3.2 Texture of igneous rocks

What texture might we expect an igneous rock to have? An igneous rock will contain crystals that grew as the magma cooled. Each crystal will have started to grow unhindered by neighbouring crystals, so an igneous rock therefore has a crystalline texture in which the crystals are randomly oriented.

To picture this, consider a magma, at an initial temperature of perhaps 1000 °C, as it slowly cools underground (Figure 2.3, path (a) to (d)). Initially the magma is completely molten (Figure 2.3a) but, unlike water placed in the ice box of a freezer, magma doesn't turn from being totally liquid to totally solid at a single temperature when it is cooled. Instead, different minerals crystallize over a range of temperatures (in fact over one or two hundred degrees Celsius). So, when the temperature of a magma falls by a small amount, only a few mineral crystals will form (Figure 2.3b). On further cooling these crystals grow larger, and new minerals also start to crystallize (Figure 2.3c). Eventually, these crystals form an interlocking network, with the last crystals to grow filling the spaces between. When totally solidified, the rock has the crystalline texture shown in Figure 2.3d.

In contrast, very fast cooling allows crystallization to occur by the nucleation of many small crystals rather than the steady growth of a few crystals. The resultant igneous rock contains innumerable tiny crystals that may be so tiny as to be indistinguishable except under the high magnification of a microscope (Figure 2.3, path (a) to (e)). In the most extreme case crystallization is totally inhibited and the starting liquid is quenched to form volcanic glass.

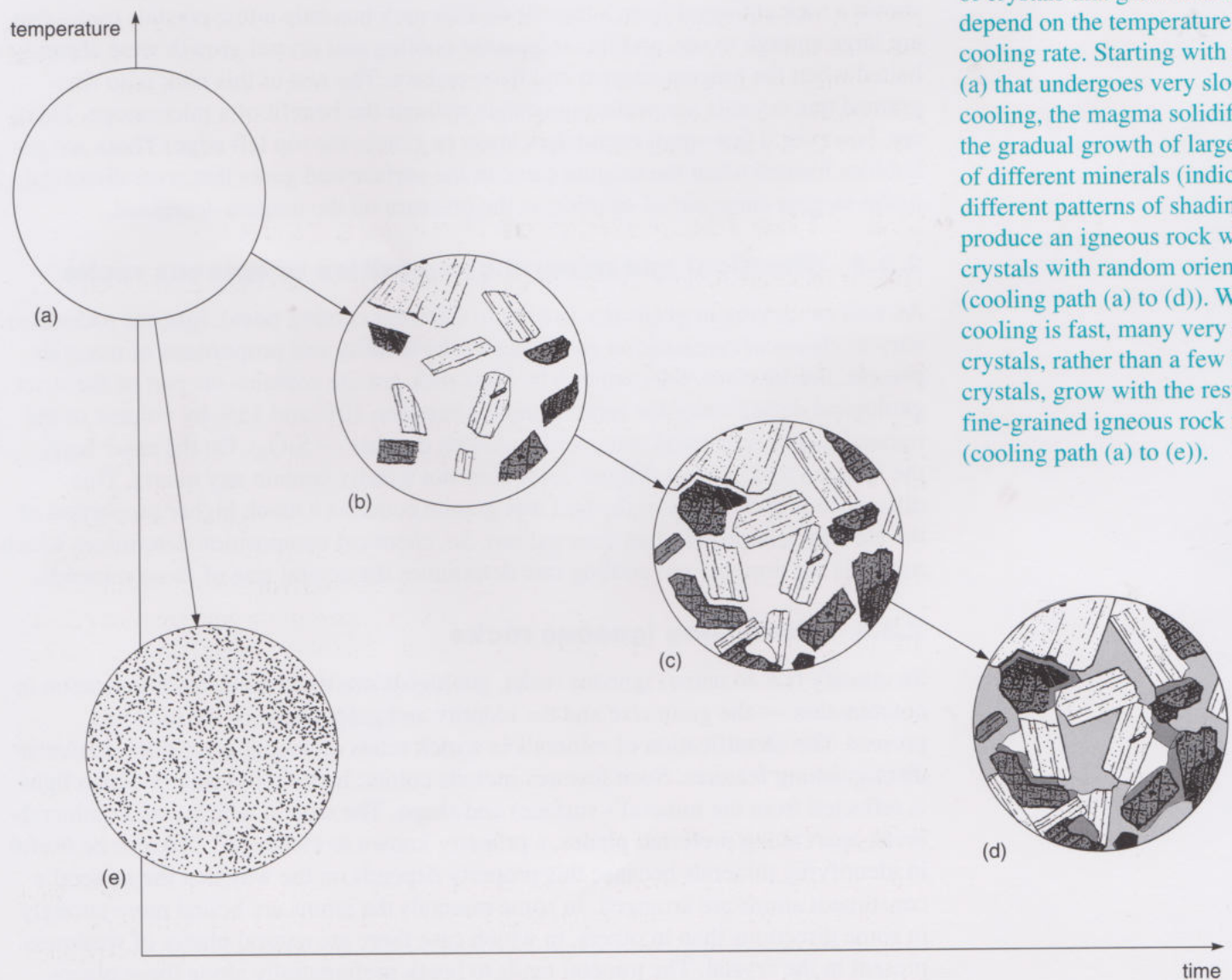


Figure 2.3 The number and size of crystals that grow in a magma depend on the temperature and the cooling rate. Starting with a liquid (a) that undergoes very slow cooling, the magma solidifies by the gradual growth of large crystals of different minerals (indicated by different patterns of shading) to produce an igneous rock with large crystals with random orientations (cooling path (a) to (d)). When cooling is fast, many very tiny crystals, rather than a few large crystals, grow with the result that a fine-grained igneous rock is formed (cooling path (a) to (e)).

- Would you expect a fine-grained igneous rock to have formed deep below the Earth's surface or at the surface?
- A fine-grained igneous rock requires rapid cooling and this is more likely at the surface, where magma comes into contact with air or water, rather than in the hot interior of the Earth.

Generally speaking, the number and size of the crystals in an igneous rock depend on the amount of time available for their growth. The slower the cooling, the bigger the crystals. In the case of extrusive rocks, the amount of time is short — anything from a few seconds for droplets of magma flying through the air in an explosive volcanic eruption, to a few years for the interior of a thick lava flow. This results in small crystals (Figure 2.3e). For intrusive rocks, the cooling rate is much slower and there is time for larger crystals to grow (Figure 2.3d). (The times involved are not known for certain because the magma at depth cannot be observed.) Figure 2.4 illustrates this with two rocks of essentially similar chemical composition, but from different igneous settings. Figure 2.4a shows an intrusive rock containing crystals of mainly two minerals — one dark (this is the iron (symbol Fe) and magnesium (Mg) -bearing silicate mineral called pyroxene), the other pale (this is the calcium (Ca), sodium (Na), aluminium (Al) -bearing silicate mineral called plagioclase feldspar). The crystals are intergrown (cf. Figure 2.3d) and easily visible. In contrast, Figure 2.4b shows a rock collected from a lava flow. This rock has only a few crystals (pale) that are large enough to see, and this is because cooling and crystal growth were abruptly halted when the magma erupted and froze as lava. The rest of this rock is so fine-grained that crystals are indistinguishable without the benefit of a microscope. There are, however, a few small round dark areas (e.g. near the top left edge) These are gas bubbles formed when the magma came to the surface and gases that were dissolved in the magma came out of solution as the pressure on the magma decreased.

2.3.3 Chemical and mineral composition of igneous rocks

As well as varying in *grain size* (owing to different cooling rates), igneous rocks also vary in *chemical composition* and hence in the identity and proportions of minerals present. For instance, the common igneous rock granite contains (as part of the strict geological definition of the term 'granite') between 10% and 35% by volume of the mineral quartz (chemical composition silicon dioxide — SiO_2). On the other hand, the igneous rock gabbro (Figure 2.4a) does not usually contain any quartz. This difference is due simply to the fact that granite contains a much higher proportion of the element silicon (Si) than does gabbro. So, chemical composition determines which minerals are present, and cooling rate determines the crystal size of those minerals.

2.3.4 Classifying igneous rocks

To classify (i.e. to name) igneous rocks, geologists use three pieces of information in combination — the grain size and the identity and proportions of the minerals present. The identification of minerals in a rock relies on recognizing their particular distinguishing features. Such features include colour, lustre^G (the way in which light is reflected from the mineral's surface) and shape. The way in which certain minerals break apart along preferred planes, a property known as cleavage^G, can also be useful in identifying minerals because this property depends on the way that the mineral's constituent atoms are arranged. In some minerals the atoms are bound more strongly in some directions than in others, in which case there are natural planes of weakness present in the crystal. The mineral tends to break preferentially along these planes.

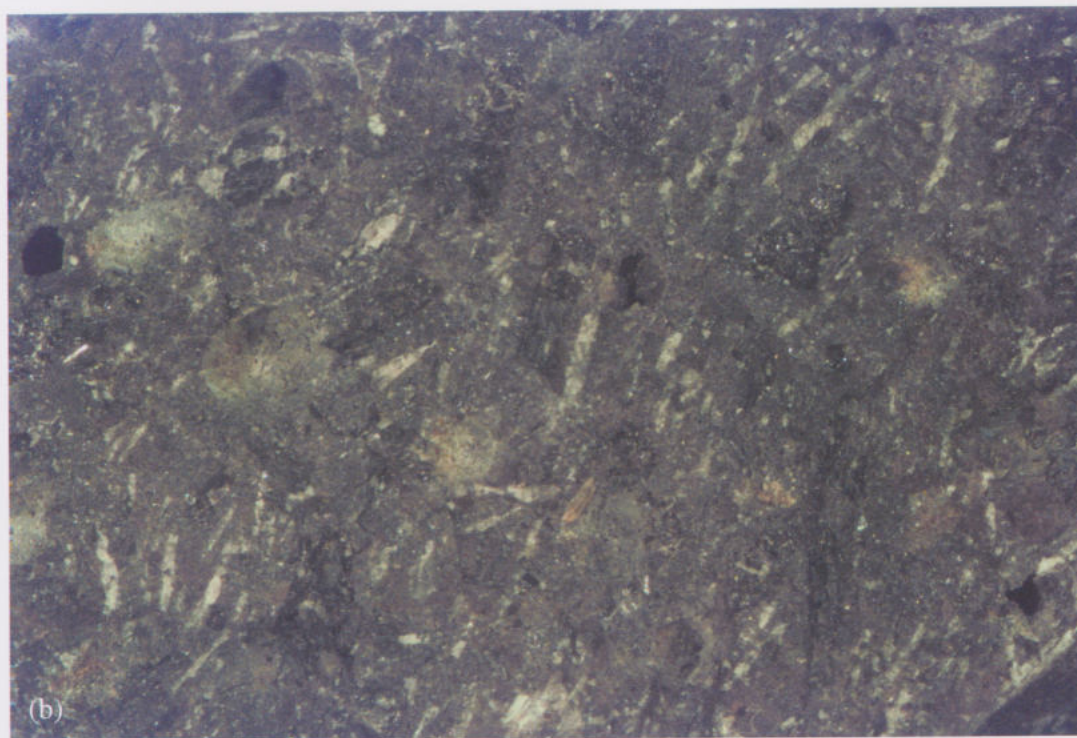
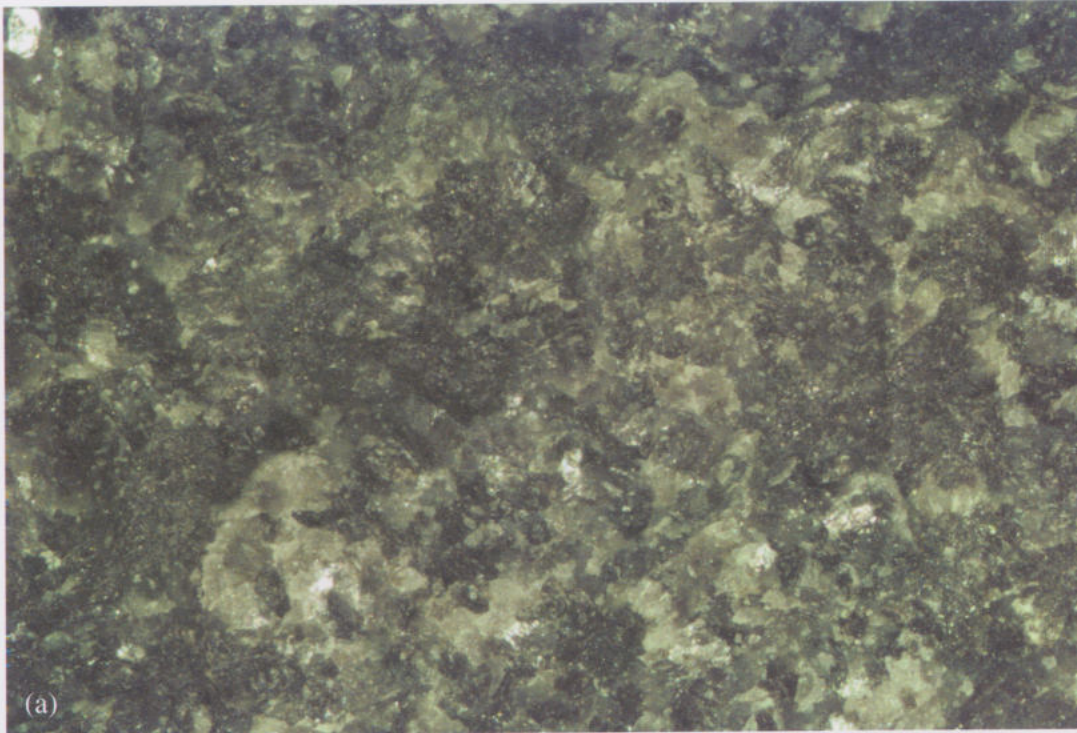
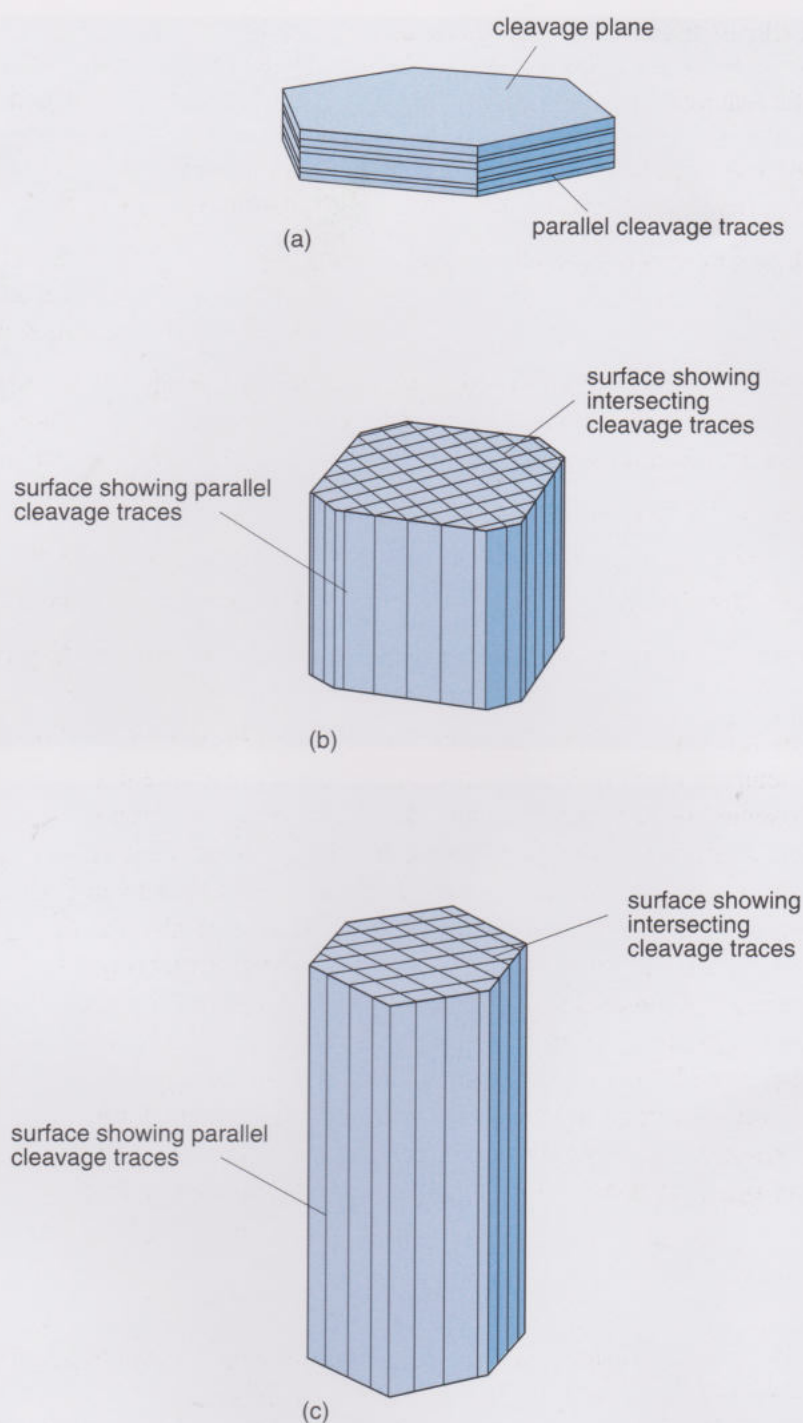


Figure 2.4 Close-up views of two rocks formed by the cooling of magma of the same chemical composition. (a) Gabbro formed by slow cooling in a subterranean magma chamber. (b) Basalt formed by rapid cooling of magma at the Earth's surface in a lava flow. The field of view is about 2.4 cm across in each case.

One mineral that shows this feature clearly is mica, a silicate mineral containing potassium (K), iron (Fe), magnesium (Mg) and aluminium (Al), together with hydroxyl (OH) groups. These compositional details needn't distract us, but the important point is that the potassium ions occur in layers, which separate sheets of more tightly bound silicon, oxygen and other atoms. Mica therefore splits apart parallel to the sheets, so has just one set of cleavage surfaces parallel to each other (Figure 2.5a). This is why mica forms platy or flake-like crystals, rather like the pages of a book. Other minerals can have two or three sets of cleavages (which intersect at characteristic angles) whereas others, notably quartz, have no cleavage and break irregularly.

Figure 2.5 Some examples of cleavage shown by different minerals. Note that in each case the number of cleavage traces with different directions depends on which face of the crystal is being looked at. (a) Mica, one set of cleavages when viewed edge on. (b) Pyroxene, two sets of cleavages intersecting at about 90° when viewed end on. (c) Amphibole, two sets of cleavages intersecting at about 120° when viewed end on. Table 1 gives additional information about these minerals, including their true colours.



The most common silicate minerals found in igneous rocks (most of which are also common in sedimentary and metamorphic rocks) are listed in Table 1: minerals rich in silicon appear towards the top, and those poorer in silicon but richer in magnesium and iron appear towards the bottom of Table 1.

- Would you expect an igneous rock that contained a lot of magnesium (Mg) and iron (Fe), but not much silicon (Si), to contain all of the minerals in Table 1 in equal abundance?
- No. The type and proportions of the minerals present must reflect the chemical composition of the rock. So an igneous rock that contains a relatively high proportion of Mg and Fe and a low amount of Si would be expected to contain more olivine and pyroxene and very little, if any, quartz.

Table 1 The common silicate minerals of igneous rocks.

Mineral	Diagnostic features	Chemical composition
Quartz	Colourless to pale; glassy lustre (broken surfaces similar to broken glass); curved fractures; no visible internal structure	SiO ₂
Feldspar ^G	Pale pink or white; two good cleavages	Ca, Na, Al silicates. Feldspar that is rich in Ca and Na is known as plagioclase, if rich in K and Na, alkali feldspar.
Mica ^G	Forms platy crystals that split into flakes; one perfect cleavage. Brown or black mica is called biotite; colourless or silvery brown mica is called muscovite	K, Mg, Fe, Al silicate containing OH. Biotite is darker than muscovite because it contains more Mg and Fe.
Amphibole	Dark green or black; elongate crystals with two cleavages (often hard to see)	Na, Ca, Mg, Fe, Al silicate with hydroxyl group
Pyroxene ^G	Dark green, dark brown or almost black; two good cleavages	Fe, Mg, Ca silicate
Olivine ^G	Grass green; no cleavage, but has irregular fractures	Mg, Fe silicate

The variation in the mineral content of igneous rocks is shown in Figure 2.6, and this diagram provides the means to classify igneous rocks. The percentage of each mineral present is represented on the vertical scale, and the range of rock type (effectively the chemical composition) is given on the horizontal scale. Coarse-grained rocks (typical grain size greater than 2 mm) are named separately from fine-grained rocks (typical grain size less than 0.25 mm). Notice that the proportion of pale-coloured minerals (felsic minerals) increases from left to right, reflecting the increase in the concentration of silicon in the chemical composition of the rocks. According to the diagram, granite is defined as a coarse-grained igneous rock containing quartz, feldspar (both alkali and plagioclase feldspar), mica and sometimes amphibole. The relative proportions of these different minerals are given by the width of the appropriate band on the diagram. For instance the quartz content of granite can range from about 10% to 35%, and the mica content from about 5% to 15%.

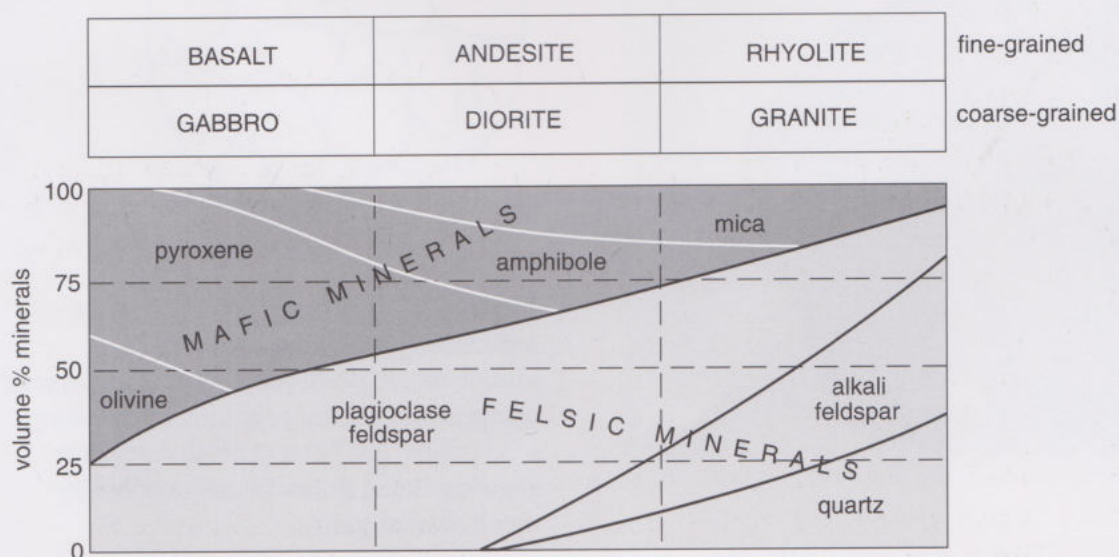


Figure 2.6 Classification of common coarse- and fine-grained igneous rocks according to their mineral content. (Medium-grained rocks are omitted for simplicity.) Dark-coloured minerals are known as mafic minerals (they are rich in magnesium and iron (Fe)). Pale minerals are known as felsic minerals (from *feldspar* and *silicon*).

Question 2.1

Do igneous rocks with small crystals form at the surface of the Earth?

Give reasons for your answer. ◀

Question 2.2 Use Figure 2.6 to decide which of the following statements are true.

- A Olivine is not present in granite.
- B The proportion of pyroxene in gabbro can vary from about 25% to 50%.
- C Alkali feldspar is less abundant than plagioclase feldspar in coarse-grained igneous rocks.
- D A coarse-grained igneous rock containing 60% plagioclase, 15% pyroxene, 15% amphibole and 10% mica is classified as diorite.
- E A coarse-grained igneous rock containing 50% plagioclase, 35% pyroxene and 15% amphibole is classified as gabbro.
- F Amphibole can be present in gabbro, diorite and granite. ◀

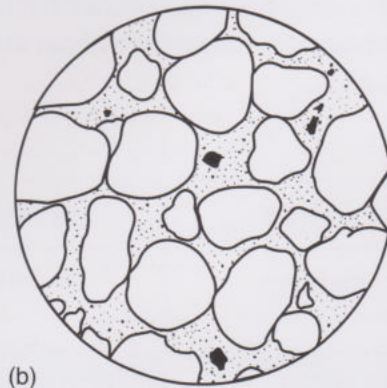
2.4 The formation of sedimentary rocks

2.4.1 Sedimentary material

The laying down, or deposition, of layers of rock fragments, mineral grains, or biological material, such as the shells or other hard parts of dead organisms, can produce sedimentary rocks. Once deposited, the loose, unconsolidated sediment may be converted into a solid rock by compaction and cementing of the grains together by chemical action deep below the surface. These rocks consist, therefore, of fragments of sedimentary material, bound together by even smaller fragments, or some sort of cementing material, and so display a fragmental texture, as shown in Figure 2.7.



(a)



(b)

Figure 2.7 The fragmental texture of sedimentary rocks illustrated by (a) a sample of conglomerate, about 20 cm across, comprising rounded pebbles and enclosing sand grains, and (b) a generalized sketch showing round grains bound together by much smaller grains.

2.4.2 Sedimentary processes

Sedimentary grains are formed when the rocks at the Earth's surface are slowly broken up physically by exposure to wind and frost, and decomposed (chemically) by rainwater or biological action. These processes are collectively termed weathering^G. Once a rock has been broken up by weathering, the small rock fragments and individual mineral grains can be eroded from their place of origin by water, wind or glaciers and transported to be deposited elsewhere as roughly horizontal layers of sediment. The resulting sediment reflects the original rock types that were weathered, the efficiency of erosion and transport, the extents of chemical and physical degradation of the sediment grains during transport, and the conditions under which the grains were deposited from the transporting water, wind or ice. For example, sand-sized grains of quartz are one of the main constituents of sandstone, but those grains may have been transported by water in a river, carried by waves on a sea-shore, or blown around in hot desert sandstorms (to give just three possibilities). How might we distinguish which of the many possibilities is the most likely in any given case?

One approach is to use the size and shape of the grains in a sediment or sedimentary rock to reveal quite a lot about the origin of the sediment. For example, a vigorous river transports much larger grains than a gentle current in a lake, so the size of the grains gives an indication of the strength of the currents that could have transported and deposited the grains. In other words, the grain size depends on the energy of the environment in which the sediment was deposited. The general shape of the grains will tell you about the nature of the transporting medium; for example, was it water or air? (See Box 2.2, *A story in a grain of sand*).

- Are all sediments composed of fragments of rock and minerals eroded from pre-existing rocks?
- No, some sedimentary rocks also contain the remains of dead organisms, i.e., fossils of plants or animals that were living at the time the sedimentary material was deposited.

Any record of ancient life preserved in a rock is known as a fossil^G - sometimes fossils are rare, whereas a few rocks are composed of virtually nothing else but fossils. In particular, many limestones were formed by the accumulation of the calcite (calcium carbonate — CaCO_3) shells and skeletons of certain marine organisms. Chalk^G is a well-known type of limestone that outcrops extensively across southern England; it is almost pure calcite, and consists largely of minute calcite plates of countless planktonic algae (phytoplankton^G) fossils. Other limestones, such as those found in the Peak District of northern England, contain abundant fossils of reef-building corals. Another example of a biologically formed sedimentary rock is coal, which is formed from compressed layers of woody plants.

Fossils are important when reconstructing the geological past because they are records of the environment at the time and place the fossil organisms were living. For instance, limestones rich in corals typically indicate warm shallow seas — the conditions needed for a coral reef ecosystem to thrive.

It is important that as many lines of evidence as possible are used to give a consistent interpretation of a rock's origin. No single feature should be taken as unequivocally diagnostic. For instance, think of a desert sand, composed of well-rounded, red oxide-coated, well-sorted quartz grains. Now imagine that climatic conditions change and these sand grains are swept away by flowing rivers and re-deposited elsewhere. The

Box 2.2 A story in a grain of sand

Quartz is a hard mineral that is very common in the Earth's continental crust. It is resistant to attack by chemicals and physically strong, so it tends to survive the weathering process that disaggregates and decomposes pre-existing rocks. Many sedimentary rocks contain grains of quartz. Quartz grains are recognizable by their glassy appearance (particularly on freshly broken surfaces) and lack of cleavage. Quartz is also hard enough to scratch steel.

Whether quartz grains are transported and deposited by moving air (by being blown around by the winds in a desert) or by moving water (in a river or in ocean currents) determines how rounded they become (Figure 2.8). The degree of rounding of quartz grains depends on the intensity and frequency with which grains collide with each other, and these factors depend on the environmental conditions. Air is less viscous than water, so windblown quartz grains collide more violently than quartz grains carried in water, which has a cushioning effect. Also, the wind speeds needed to move a sand grain of given size are higher than the speeds for flowing water. This means that collisions between grains will be much more energetic in air than in water, so the corners of windblown grains are readily knocked off, and the grains are usually very much more rounded (Figure 2.8b) than water-transported grains (Figure 2.8a).

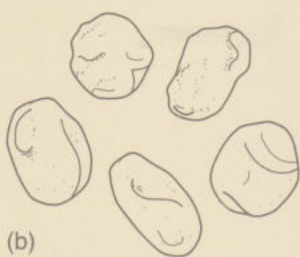
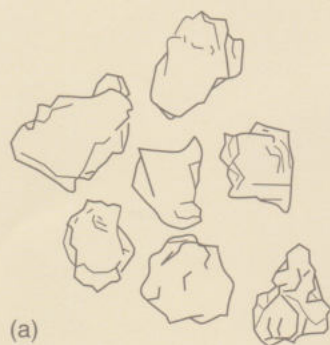


Figure 2.8 Examples of variation in grain shape associated with difference in sedimentary environment:
(a) river sand ($\times 20$);
(b) wind-blown sand ($\times 20$).

Wind-blown sand deposited in desert environments differs from water-lain sand in another way. Quartz sand grains in a desert often have a coating of red or orange iron oxide. This red-orange coloration is typical of desert landscapes and it is derived from the insoluble 'rusty' residue from weathering of iron-rich minerals. Water-lain sand grains lack such an obvious coating.

The degree of sorting in a sediment is another useful method for distinguishing different types of depositional situation. Sorting^G is a measure of the range of grain sizes present in a sediment or sedimentary rock. A poorly-sorted sediment (Figure 2.9a) has a wide range of grain sizes as a result of rapid deposition, such as occurs during a storm. On the other hand, a well-sorted sediment has a narrow range of grain sizes (Figure 2.9c), and is the result of extensive reworking of a sediment by wind action in deserts, or wave action on beaches and in shallow shelf seas.

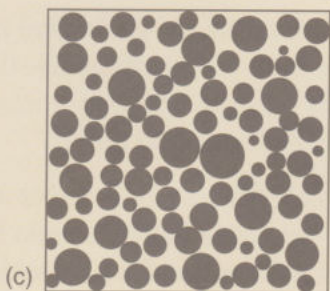
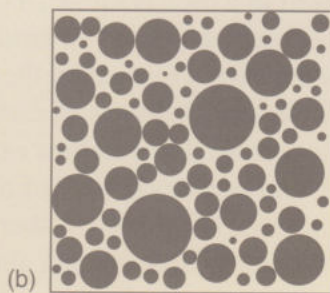
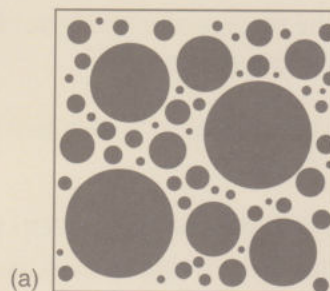


Figure 2.9 (a) Poorly-sorted sediment. (b) Moderately-sorted sediment. (c) Well-sorted sediment.

new sand deposit will be produced by the action of flowing water, but the sand grains may retain many of the characteristics of wind-deposition inherited from their previous history. A misleading interpretation can be reached if other lines of evidence are ignored. Such supplementary evidence could come from any fossils in the rock and the nature of adjacent sedimentary layers. Getting all of the necessary information involves a mixture of making observations and measurements at rock exposures in the field as well as examination and analysis of samples in the laboratory.

2.4.3 Sedimentary strata

We've seen that the detective work of piecing together a part of Earth's history from sedimentary rocks involves detailed investigation of rock samples, but this can give only a partial picture. On the larger scale of a rock exposure, there can be plenty for us to see and to interpret. As you'll see at Residential School, sedimentary rocks are usually found as layers referred to as strata^G (Figure 2.10), with each stratum (layer) recording the particular conditions at the time of its deposition. (Note: sedimentary layers are often referred to as beds, but strictly speaking the term 'bed^G' is reserved for those strata thicker than 1 cm; thinner layers are known as laminae (singular lamina).) Over time, conditions may have changed, either gradually or quickly, causing the nature of the sediment being deposited to change. In this way, a vertical stack of sedimentary strata is a record of changing conditions during a segment of geological time. The oldest sediments will be at the bottom, with progressively younger strata laid down on top. Geologists refer to this as the principle of superposition^G — older rocks are overlain by younger rocks; an individual layer is younger than the one beneath it and older than the one above it; the oldest layer lies at the bottom. This provides a relative time-scale. Changes in sedimentary rocks (or in the types of fossil they may contain) up through a sequence of strata provide a record of changing conditions over the passage of time throughout which the rocks were deposited.



Figure 2.10 A succession of sedimentary strata (layers) exposed in a cliff on the Dorset coast.

2.4.4 Stratigraphy and geological time

Stratigraphy^G is the study of how the types of strata have varied over time and also considers how they are distributed geographically. One of the most useful results of stratigraphy is a generalized geological succession — the stratigraphic column^G — that defines the divisions of geological time. Figure 2.11 shows the geological time-scale. The broadest division of Earth history is into two intervals (called eons) of very different length: the Cryptozoic Eon and the Phanerozoic Eon. The Cryptozoic is a vast amount of time — from the origin of the Earth, 4 600 million years (Ma, short for mega-annum) ago, to the start of the Phanerozoic, 545 Ma ago. Cryptozoic is derived from Greek words meaning ‘hidden life’. This is the period of time when organisms did not possess hard parts such as shells, so their remains are difficult to find in sedimentary rocks. In contrast, Phanerozoic is derived from Greek words meaning ‘visible life’, reflecting the great abundance of fossils derived from the hard parts of organisms throughout this eon. The Phanerozoic Eon is divided into three *eras*^G — the Palaeozoic, Mesozoic and Cenozoic Eras (meaning ‘ancient life’, ‘middle life’, and ‘recent life’, respectively). Each of these eras is divided into a number of periods^G of unequal length (Figure 2.11). The current period, which started 1.8 Ma ago, is the Quaternary Period.

Looking at this column, you might think that there is a complete sequence of rocks everywhere, but there is not, in the same way that there are no historical records for certain times from certain areas. In the field, the Earth scientist might find rocks of, say, the Triassic Period lying above and in direct contact with rocks of the Carboniferous Period, so that rock evidence for the Permian Period is not present. This means that either there was no deposition of sediment during the Permian Period, or that these rocks were deposited, and then eroded, before rocks of the Triassic Period were laid down. This type of relationship in the geological rock record, where at a contact there are beds of the intervening age missing, is called an unconformity^G. The recognition of unconformities in the field is a key part of unravelling the geological history of an area because they represent some sort of hiatus in the conditions at the Earth’s surface where the sediments were deposited.

2.4.5 Fossils and ancient environments

An essential component of any environment is the plant and animal life that is adapted to the prevailing conditions. Fossil plants and animals are therefore wonderful sources of information about ancient environments. Plants can leave behind remains ranging from roots, leaves and twigs to seeds and pollen. Leaves and twigs are relatively fragile, and require a comparatively low energy environment (e.g., the mudflats of an estuary) for their preservation. Seeds, pollen and spores are surprisingly robust, and are often the only parts of a plant to survive. Animals can be preserved in one of two ways – either by some part of their body remaining as a fossil, or by some trace, such as the animal’s footprints preserved in a muddy sediment, becoming preserved as sedimentary rock (a trace fossil). Even dung can end up fossilized, the resultant fossils being known as coprolites.

Body fossils of animals include shells, skeletal frameworks, bones and teeth. We do not have room here to describe the most common fossil groups; you will be able to find examples during your Residential School that can be discussed in your group with your tutor. Figure 2.12 illustrates some of them, with their characteristic features labelled. It is more important that here we consider how an organism gets fossilized, and what they can tell us about the environment in which they lived.

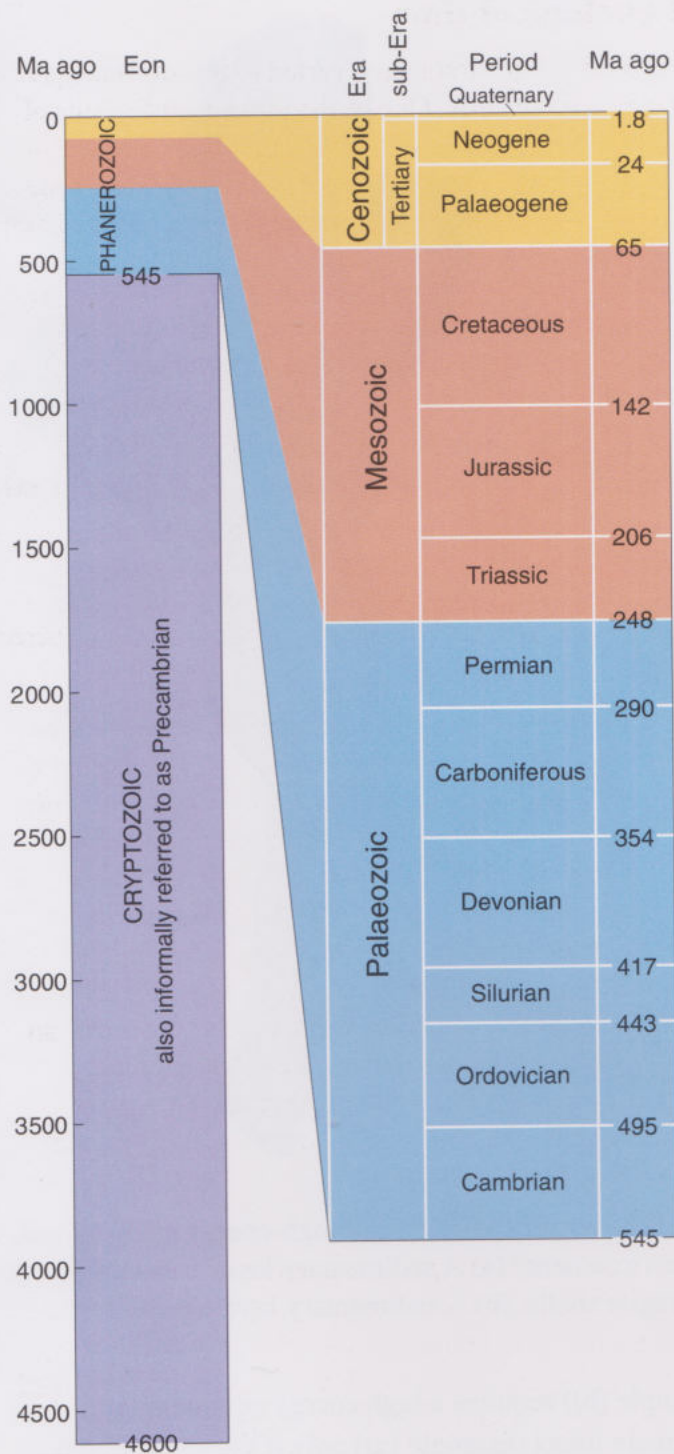


Figure 2.11 The geological time-scale.

First of all, we have to consider the organism itself. Does it have any hard parts? If not, e.g. a jellyfish, then its chances of fossilization are very low indeed. However, some soft-bodied organisms burrow into sediment, e.g. the lugworm that leaves the familiar worm casts on the shore, and at least there is the chance that their burrow becomes preserved as a trace fossil. An organism with a one-piece shell, such as a periwinkle or garden snail, stands a better chance of being preserved intact, than if its skeleton is made up of lots of pieces, like a sea urchin or crinoid (Figure 2.12c).

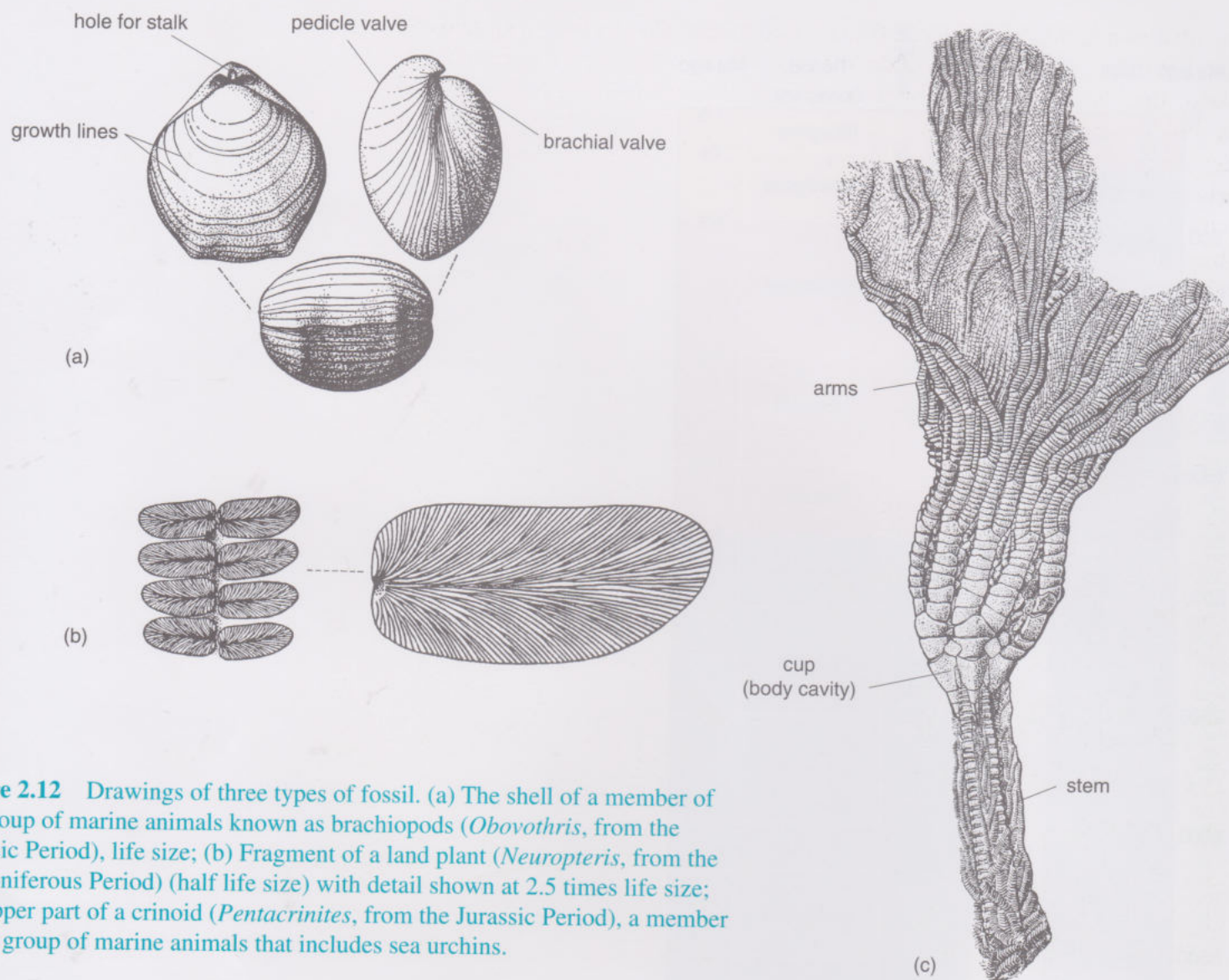


Figure 2.12 Drawings of three types of fossil. (a) The shell of a member of the group of marine animals known as brachiopods (*Obovothis*, from the Jurassic Period), life size; (b) Fragment of a land plant (*Neuropteris*, from the Carboniferous Period) (half life size) with detail shown at 2.5 times life size; (c) Upper part of a crinoid (*Pentacrinites*, from the Jurassic Period), a member of the group of marine animals that includes sea urchins.

- Which of the following is indicative of deposition in a high-energy environment, and which in a low-energy environment? (a) A sedimentary layer containing complete examples of thin fragile shells. (b) A sedimentary layer containing broken shell fragments.
- For shells to be broken (example (b)) requires a high energy environment, whereas fragile shells can remain intact (example (a)) only if the speed of the transporting and depositing currents are low (i.e., a low energy environment).

Fossils of land-dwelling organisms (e.g., Figure 2.12b) indicate deposition in or at least near a terrestrial environment, whereas marine fossils indicate a marine environment suitable for the fossil organisms to have lived in (e.g., correct temperature range, light levels, salinity, water depth). This evidence can then be added to evidence from the features of sedimentary grains (Section 2.4.2) to help reconstruct the environment.

2.4.6 Sedimentary structures

Consider some of the places where sedimentary materials are moved and deposited. Are the sediments always laid down in perfectly horizontal, perfectly flat layers? No;

as often as not, the depositing surface is not perfectly flat. Instead, a system of parallel ridges, or ripple marks, like the ones shown in Figure 2.13a, form by the action that flowing water has on the erosion, transport and deposition of sand grains. In sedimentary strata, ripples and dunes (which are effectively larger versions of ripple marks) are normally seen in cross-section rather than in plan view, because pristine depositional surfaces are rarely exposed. In cross-section, the thin sedimentary layers that built up each ripple or dune can be visible, inclined at an angle to the top and bottom of the bed as a whole (Figure 2.13b). These structures are called cross-stratification^G and are a clue to the strength, direction and setting of the currents that produced them.



Figure 2.13 (a) Ripple marks on a sandy beach. (b) Cross-stratification in sandstone formed by underwater dunes advancing from right to left; the exposure is about 3 m high.

Many features associated with either erosion or deposition in a sedimentary environment can be preserved. For instance desert dunes can yield cross-stratification on a scale measured in many metres. All such structures are referred to as sedimentary structures and, as we have just seen, provide evidence with which to interpret ancient sedimentary environments. At Residential School you will have the opportunity to see sedimentary structures forming on beaches.

Question 2.3 Explain whether the energy of the environment of deposition of the following sediments is (i) high, (ii) medium or (iii) low.

- (a) A sediment made up of pebbles and boulders.
- (b) A sediment made up of fine clay particles.
- (c) A sediment made up of sand grains. ◀

2.5 The formation of metamorphic rocks

Any type of rock can become a metamorphic rock if it is heated to temperatures of several hundred degrees Celsius, and/or if subjected to high pressure (because of the weight of overlying rocks). During metamorphism^G, the minerals making up the rock become chemically unstable, meaning that their constituent ions are redistributed. The result is that either large crystals grow at the expense of existing smaller ones, or a new set of minerals is formed. Generally speaking, the overall chemical composition of the rock remains about the same. Although igneous and metamorphic rocks both form at high temperatures, an important distinction is that metamorphism occurs in the solid state, whereas igneous activity involves liquid rock (magma).

2.5.1 Causes of metamorphism

- What natural process could cause a rock to be heated?
- Heating can be caused when hot magma is intruded into a cool rock.

On the other hand, an increase in both pressure and temperature will come about if the rock becomes more deeply buried as a result of Earth movements, particularly at convergent plate boundaries where continents collide, or is covered by a deepening layer of sedimentary deposits.

There are therefore two settings where metamorphic rocks can be found. In the first, and simplest situation, a narrow zone around the edge of an igneous intrusion becomes heated by the magma and undergoes metamorphic recrystallization. This is known as contact metamorphism because it is caused by hot magma coming in to contact with cold rocks; contact metamorphism is due to heating alone.

The second common setting of metamorphism is far more extensive, and is caused by the deep burial of crust at continental collision zones. In this setting, huge volumes of rock experience increases in temperature and pressure, causing metamorphism on a regional scale; this is known as regional metamorphism.

2.5.2 Metamorphic recrystallization

To consider metamorphic recrystallization at its simplest, let's begin by imagining a sedimentary rock composed entirely of quartz grains — a quartz sandstone. Sandstone is a sedimentary rock and so has a fragmental texture (see Figure 2.7b). When it is subjected to high temperature and high pressure no new minerals can form

because there are no other minerals present with which the quartz grains could react. All that can happen is that the quartz grains recrystallize and the rock known as quartzite is formed. The original fragmental texture is obliterated and replaced by a crystalline texture. Likewise, when a pure limestone, comprising calcite (i.e., CaCO_3) is metamorphosed, the calcite recrystallizes, and marble is produced. In these cases the rock adjusts to a high pressure or temperature by slowly recrystallizing in a denser, more compact, form.

In most cases, however, we start with a rock containing several different minerals, giving a richer chemical mix for metamorphic reactions to work with. At sufficiently high temperature or pressure, the original minerals react with each other, and new mineral crystals grow. In order to do this the constituent atoms must diffuse at different rates through the rock, but diffusion is extremely slow so atoms can move only very small distances in a given amount of time. The chemical rearrangement of the rock therefore entails the growth of small crystals unless the temperature is particularly high, in which case larger crystals can grow. Whenever metamorphism occurs in the compressional environment of continental collision zones (regional metamorphism), the rock is also subject to directed pressure and this also has an effect on the way the minerals crystallize. In the case of the metamorphism of mudrocks, mica crystals are formed during metamorphism. Mica crystals are characteristically platy in shape, reflecting the fact that the atoms in mica are arranged in layers or sheets (Section 2.3.4). When platy minerals grow during metamorphism the energetically most favourable pattern of growth is one in which their flat surfaces lie more or less parallel, and at right angles to the main direction of imposed pressure (Figure 2.14). Likewise, any elongate crystals grow aligned parallel with each other.

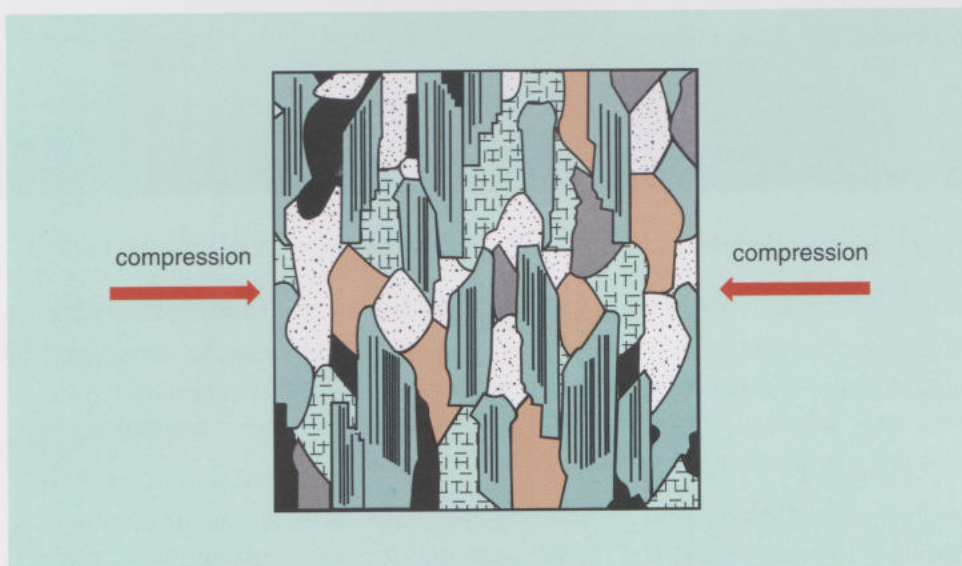


Figure 2.14 A sketch of the interlocking crystalline texture in a metamorphic rock formed by regional metamorphism; mineral banding develops at right angles to the direction of compression.

- In what way is the metamorphic texture shown in Figure 2.14 the same as the texture found in igneous rocks (Figure 2.3) and in what way does it differ?
- The metamorphic rock has a crystalline texture, like igneous rocks. However, it also has mineral layering, or alignment, whereas the crystals in igneous rocks have grown in random directions.

The alignment of platy minerals means that the rock has a series of closely spaced planes of weakness running through it, so it tends to split apart into fairly flat pieces. The rock is said to have a foliation^G, because its mineral grains are aligned like a

stack of leaves (foliage) lying one on top of the other. A classic example is slate^G, a metamorphic rock with an extremely fine grain size that can be split into thin sheets used for roofing. Slate, and the coarser-grained metamorphic rock schist^G (pronounced 'shh-ist', to rhyme with 'mist') (Figure 2.15) have marked foliation because they contain a lot of mica. But not all metamorphic rocks are as rich in mica. In such cases metamorphism can still produce a banding effect, but here the minerals grow segregated into alternating bands a few millimetres to centimetres thick. Bands of light-coloured minerals alternate with bands of dark-coloured minerals, and the rock is known as gneiss^G (pronounced 'nice'). In general, the coarser the grain size of a metamorphic rock, the higher the temperature and/or pressure.

Figure 2.15 The metamorphic rock schist. The platy mica crystals (dark) are arranged in undulating layers, defining a foliation. The sample is about 6 cm long.



- Is the control on grain size in metamorphic rocks the same as in igneous rocks?
- No. In igneous rocks grain size is controlled by cooling rate (and crystals grow from a liquid); the slower the cooling, the coarser the grain size. In metamorphic rocks grain size is controlled by pressure and temperature (and crystals grow by transformation of existing minerals in the solid state); the higher the temperature and pressure, the coarser the grain size.

The end product of metamorphism depends on two main variables — the chemical composition of the starting rock, and the pressure and temperature conditions under which metamorphism occurred. A useful analogy is with cooking — the product of baking in an oven depends on the ingredients that went in (chemical composition) and the temperature that the ingredients were subjected to in the oven. Just as the shape, colour and taste of food from the oven gives us clues about the ingredients and the baking conditions, so the texture and mineralogy of a metamorphic rock allow us to say something about the original rock type and the temperature and pressure conditions in the crust where metamorphism occurred.

Question 2.4 Compare and contrast contact metamorphism and regional metamorphism in terms of (a) the presence of foliation and (b) the distribution of the affected rocks. ◀

2.6 Interlude

Now that we have covered the features found in igneous, sedimentary and metamorphic rocks, and seen how these features can be explained by the processes that formed the rocks, here is a useful point at which to have a break before continuing with the next section. Before returning, you might like to see for yourself what types of rock you can find in your area. Can you identify their texture, or spot any fossils? Surfaces that haven't been obscured by grime or lichens are by far the best, as it is the attributes of the rock itself, not any later weathering processes that we are interested in. Polished work surfaces, ornamental stones, grave stones and shop fronts are interesting possibilities to start with. And avoid rocks whose grain size is so small that you can't recognize individual grains; a microscope is needed for studying those rocks! This is also an appropriate point at which to read *SGSG*, Chapter 8 Section 2, 'Observing', as this discusses how careful observation is one of the fundamental aspects of doing science.

Now is a suitable time to skim through Sections 1 to 3 of the Activity A workbook. You won't be able to make sense of all of it without the hand specimens of rocks. Don't worry about this — we only want you to gain some idea about what will be expected of you.

In the next section you'll do an exercise to introduce you to some of the geological skills and ideas that you'll develop in Activity C, 'Investigating the environment'.

2.7 Geological fieldwork

Although much can be learned from samples of rocks in the laboratory or at home, the 'natural habitat' of rocks is outdoors. Here the distribution and layout of different rocks is visible wherever rocks are exposed in places such as stream beds, cliffs, rocky shorelines, quarries, or road cuttings. The exposed rocks can be studied in just the same detail as individual laboratory samples, and geological fieldwork allows the size and extent of each rock unit to be seen and the relationships between them to be observed. This gives us more information from which to deduce the conditions under which the suite of exposed rocks may have formed.

It is possible that you won't have had cause to look closely at an exposure of rock before, so you will be guided through this aspect of scientific fieldwork as part of Activity C, 'Investigating the environment'. In order to deduce the conditions under which the rocks formed and the sequence of geological events that formed them, you'll be making systematic observations, and recording those observations in the 'Investigating the environment' workbook. Doing this will give you the basic information from which to work out some of the geological history of the area you'll visit.

2.7.1 Making and using field sketches

How do we start to make sense of a rock exposure? Drawing a sketch is one of the best ways to start, as it forces you to notice many aspects of the exposure. It also helps you to build up a picture of which aspects are significant and which are incidental or even irrelevant to a geological study. The aim of a field sketch is that it provides a record of your observations (along with notes taken at the same time, and also perhaps a photograph to record details). A sketch is complementary to a photograph because it allows you to highlight and label the significant features. However, your sketch need not be a work of decorative art.

Activity 2.1 Sketching a geological exposure

Now look at Figure 2.16a, a photograph of a coastal rock exposure of sedimentary strata. The rock is far from featureless, and we'll explore the visible features by first sketching the exposure, and then working through our sketched observations to develop some ideas about the history of the rocks. Figure 2.16b provides space for you to draw your sketch of the main features of the coastal exposure. Use a pencil, and follow the general guidelines 1–5 below. Spend no more than 15 minutes doing this.

- 1 Spend a minute or so looking carefully at the rock face to begin to recognize some general details.
- 2 Now it is time to start committing some of your observations to paper! Draw in the skyline and the base of the exposure. (In Figure 2.16a this extends to the bottom of the photograph.). This marks out the area that you have to concentrate on.
- 3 Sketch in the main features using simple lines wherever possible. If necessary, rub out and redraw particular areas until you're happy with them. Part of the skill of making successful geological observations of complex exposures is to follow any key feature (the top or bottom of a sedimentary bed, for instance), as far as you can with your eye until it disappears from view. Similarly, the skill of making an effective geological sketch is to draw a discrete, continuous line (where the feature is continuous), rather than to sketch a vague series of unconnected lines. Ignore shadows from sunlight and features such as loose boulders, fallen branches, etc. You can indicate any patches of vegetation using your own simple symbols for grass, trees, etc.
- 4 Label the features you have drawn.
- 5 Add a scale, the approximate orientation, for example by marking west on the left and east on the right, and a title including information identifying the location.

Once you have finished, compare your sketch with the one provided by another student in Figure 2.20 (printed at the end of Section 2), and also compare Figure 2.20 with Figure 2.16a. ◀

2.7.2 Interpretation of a geological exposure

We now want to make use of the observations obtained by sketching the exposure, and it is useful to start by briefly summarizing the features seen. First of all, you probably noticed the large boulder in the foreground of Figure 2.16a. Where did this boulder come from? It looks to have the same colour as nearby rocks so has most likely come from the same piece of coastline. However, because the boulder is not actually a part of the solid bedrock, we can't rely on it to give a true picture of the local bedrock. Boulders and any other loose rock may have been transported by natural processes or human activities, so geologists usually ignore them when investigating the rocks exposed at the surface. Indeed, much of Britain has been affected by glaciation with the result that the countryside is strewn with large boulders derived from far and wide and transported by moving ice — great for reconstructing the routes of past glaciers, but a shoal of red herrings when trying to investigate the underlying bedrock. We'll therefore ignore the boulder and concentrate on the rocks forming the solid cliffs and foreground.

The exposed rocks have sets of parallel lines running through them, and this might have been one of the first things you spotted when taking a careful look at the photograph. These lines define sedimentary layers or strata, so we are dealing with sedimentary rocks. In the rocks forming most of the vertical cliff face, the layers are

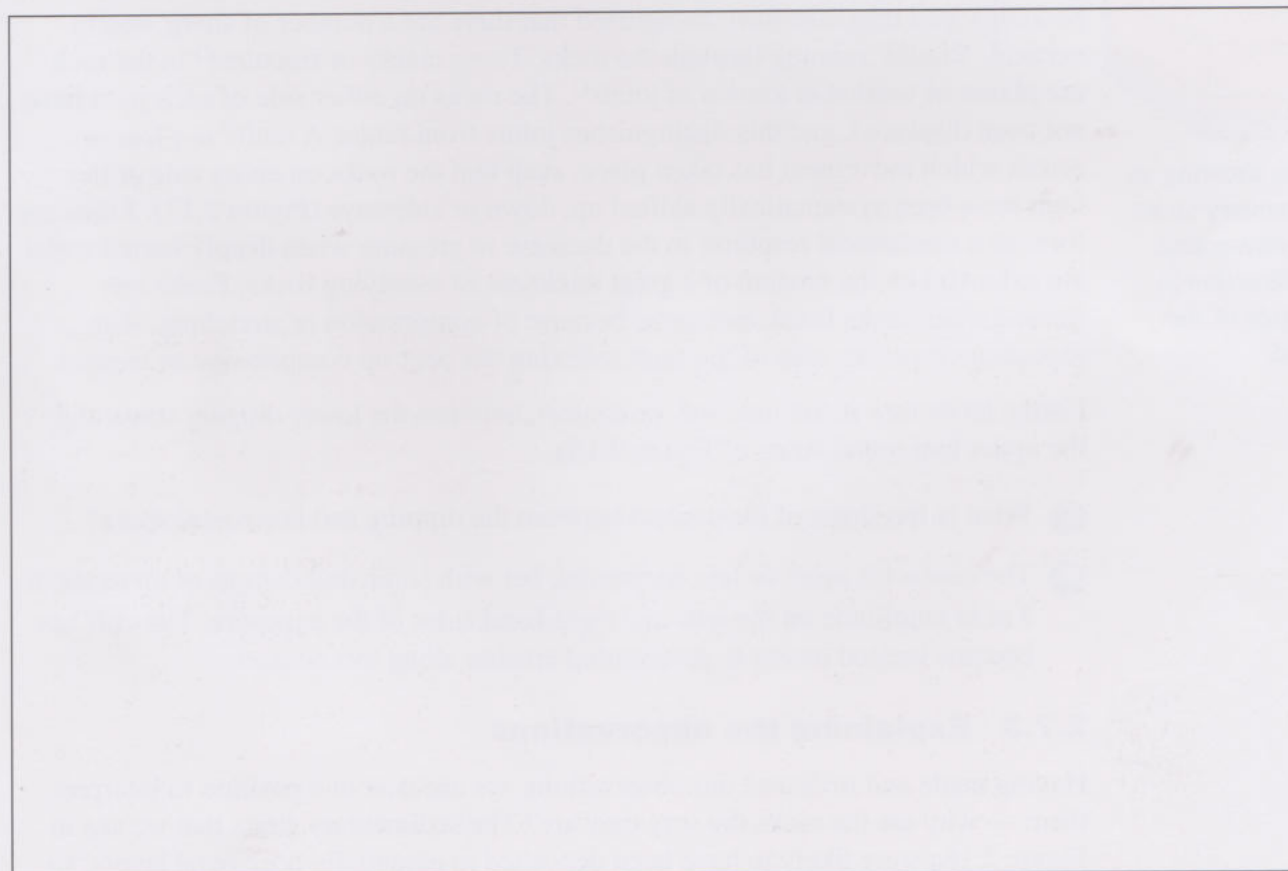


Figure 2.16 (a) A view of a coastal rock exposure about 10 m high. The view is taken looking roughly north; west is to the left and east is to the right. (b) Space for a pencil sketch of the rock exposure.

horizontal, whereas in the white rocks of the foreground the layers are inclined to the right. They are like a tilted stack of books. In this case the strata slope, or dip, to the right at an angle of about 30° . The outcrop has two distinct parts — the upper set of horizontal strata and the lower part of dipping strata.

Looking first at the dipping strata, the sediments appear fairly monotonous, with no obvious change in colour or thickness of individual beds as we pass from left to right. Close up, the rock has a pitted or pock-marked appearance in places (for example the rock face immediately to the right of the large boulder). However, this is a superficial feature caused by weathering of the rock where it is exposed to the elements. The cement holding the sedimentary grains together may not be uniform, such that weathering and erosion take place unevenly, resulting in the weakest spots becoming pockmarked. Although in itself an interesting feature, it is a detail that we won't consider further.

The horizontal strata are also pale in colour, though the vertical cliff that they form has some areas of black and pale green coloration. Rock faces that have been exposed for a long time often become discoloured, owing to chemical weathering or colonization by lichens and other organisms. In this case, the black coloration is due to biological activity, and the natural colour of the rock is visible in only a few places. Green algae are growing on the lower part of the cliff where wave action provides moisture. When working at the rock exposure an Earth scientist would carefully use a geological hammer to break open the rock and expose a fresh, unweathered, surface in order to find out the true nature of the rock. As well as spotting the fairly continuous horizontal lines defining the horizontal strata (or bedding), you may also have recognized that there are a number of sharp, nearly vertical, 'cracks' running through the rocks. These cracks or fractures^G in the rock are planes of weakness known as joints^G. The rocks on either side of each joint have not been displaced, and this distinguishes joints from faults. A fault^G is a fracture across which movement has taken place, such that the rocks on either side of the fault have been systematically shifted up, down or sideways (Figure 2.17). Joints can form as a mechanical response to the decrease in pressure when deeply buried rocks are exhumed by the erosion of a great thickness of overlying rocks. Faults are formed when rocks break and move because of compression or stretching, with movement on either side of the fault relieving the pent-up compression or tension.

Lastly, let us turn to the junction, or contact, between the lower dipping strata and the upper horizontal strata of Figure 2.16a.

- What is the shape of the contact between the dipping and horizontal strata?
- The contact is more or less horizontal, but with some undulations of up to about 1 m in amplitude on the left- and right-hand sides of the exposure. The cliff has become incised owing to preferential erosion along this contact.

2.7.3 Explaining the observations

Having made and reviewed our observations, we are now in a position to interpret them — why are the rocks the way they are? The sedimentary strata that we see in Figure 2.16a were likely to have been deposited in essentially horizontal layers, so why is one set tilted and the other horizontal? To answer this we need to think about the processes that account for each feature and the relative timings of these processes. So, the strata exposed below the cliff have undergone deposition and tilting. The strata exposed high in the cliff have undergone deposition, but have not

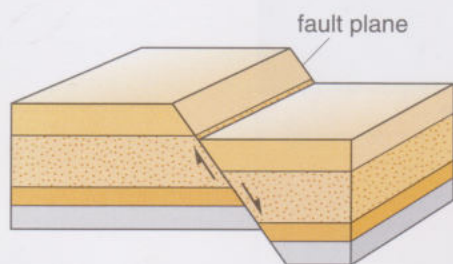


Figure 2.17 Diagram showing an example of how sedimentary strata get displaced by fault movement. The arrows show the direction in which the strata each side of the fault plane have moved.

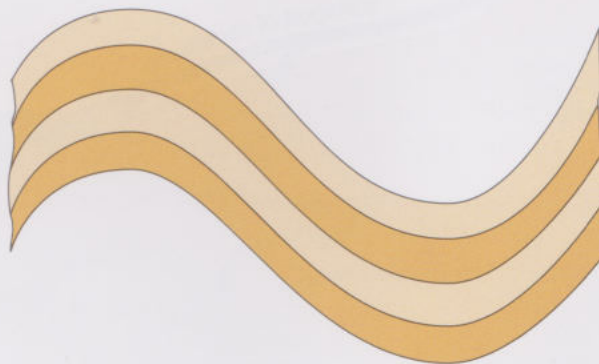
been tilted. The rocks with the more complicated history must be the older of the two sets of strata.

A sequence of events that accounts for the observations is as follows.

- 1 Deposition of sediments as horizontal layers. (Details about the environment in which these strata were deposited could be discovered by closer study of the rocks, using observations of the nature of the grains, sedimentary structure or fossils, as we saw in Section 2.4.)
 - 2 Compaction and cementation of the white sediments to form solid sedimentary rock.
 - 3 Tilting of the white strata. This could have been accomplished by powerful movements in the Earth, involving the rocks being compressed from the sides, to produce folds (Figure 2.18).
- Assuming that the rocks were not completely overturned, in which direction (left or right) would you walk to find the youngest tilted strata?
 - The youngest layers lie on top of older layers (the principle of superposition introduced in Section 2.4.3), so these will be to the far right on Figure 2.16a.



(a)



(b)

Figure 2.18 Diagram showing how initially horizontal strata (a) may become folded and hence tilted (b). The directions of the arrows in (a) show that compression is taking place.

- 4 Weathering and erosion to expose the tilted white strata at the Earth's surface, prior to deposition of the younger strata.
 - 5 Deposition of sediment onto the uneven eroded surface of tilted white strata. The hummocky contact between the two sets of strata that we spotted at the end of Section 2.7.2 is therefore the original land (or sea floor) surface when the younger strata were first deposited.
- What term is used to describe the contact between the tilted and horizontal strata?
 - It is an unconformity (Section 2.4.4).

- 6 Compaction and cementation of the younger sediments to form solid sedimentary rock.
- 7 Formation of vertical joints in the horizontal strata.
- 8 Erosion to produce the present land surface and coastal exposure.

2.8 The rock cycle

As you are reading this, rocks are being formed and destroyed on the Earth. Rocks are being heated and squeezed to form new metamorphic rocks; other rocks are melting to form magmas, which eventually cool and solidify as new igneous rocks; and the processes of weathering, erosion, transport and deposition are generating new sediments. The continuous action of rock-forming processes means that (given time) any rock in the Earth's crust will become transformed into new types of rock and that these too may in turn be transformed in to yet other rocks. We call this recycling of rock materials the rock cycle^G.

2.8.1 Moving around the rock cycle

One way of illustrating the possible ways of moving material around the rock cycle is to draw a diagram that places the processes into their geological contexts. Since the rock cycle involves processes occurring on the Earth's surface and also within its interior, we use a cross-section through the Earth's crust and uppermost mantle to do this, as shown in Figure 2.19. In this diagram we have concentrated on the most common processes within the rock cycle.

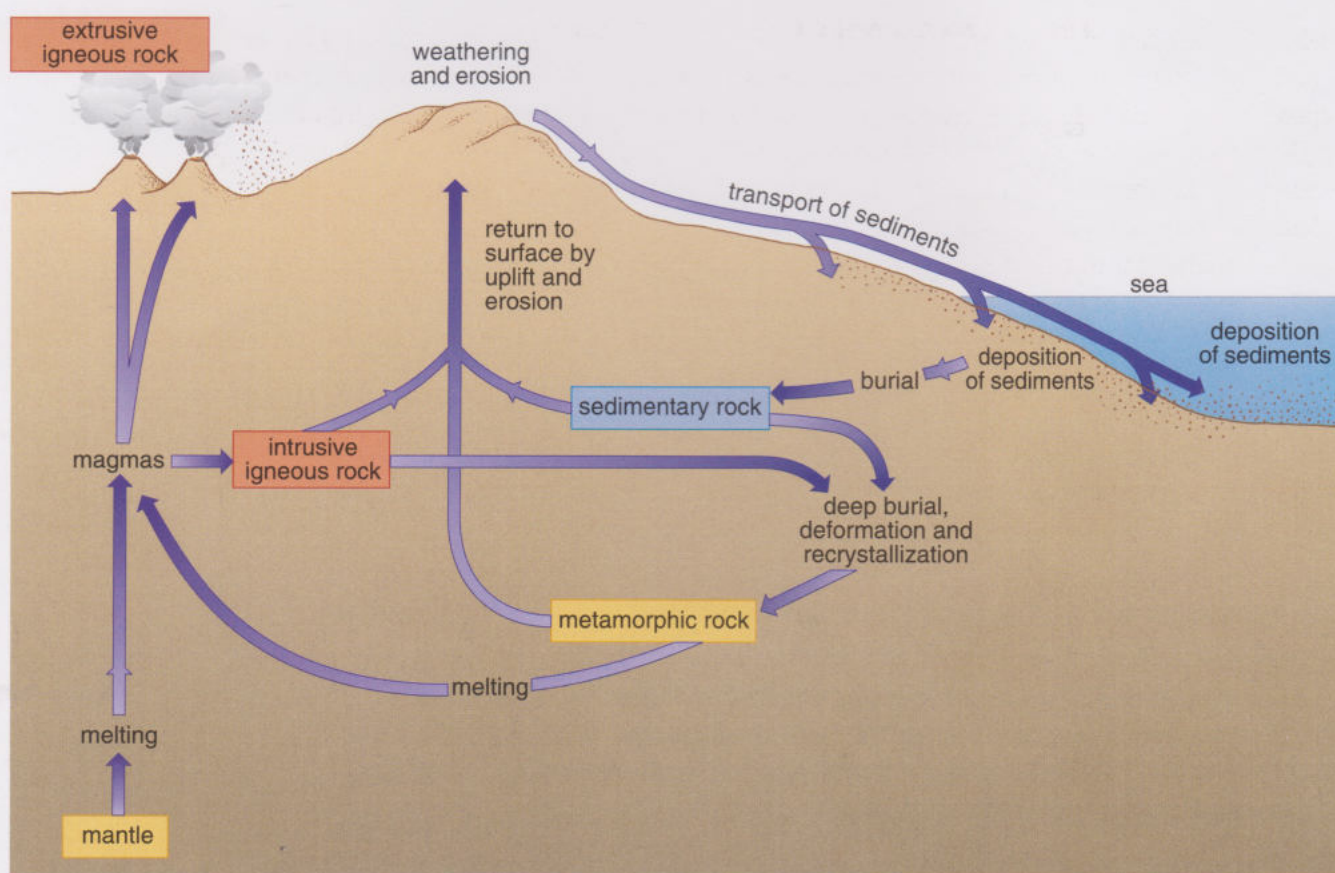


Figure 2.19 The rock cycle. The arrows show the paths and processes taken in transforming one type of rock (contained in a box) to another type.

On the grand scale of the Earth, many rock cycle processes occur at plate boundaries, where plates are created or destroyed. For example, the increases in pressure and temperature needed to cause metamorphism in the crust occur in continental collision zones. Metamorphism also occurs in the crust at subduction zones, where rocks get dragged down to great depths and hence experience great pressures. Igneous processes take place wherever melting occurs in the mantle — beneath divergent plate boundaries, above the subducting plate at convergent plate boundaries, and far from plate boundaries at hot spots.

The sedimentary processes of erosion and deposition that are important elements of the rock cycle occur in a great variety of surface environments, such as glacial environments, deserts, and continental shelves. Great rivers, such as the Nile and Mississippi, transport sedimentary material many thousands of kilometres across plate interiors, laying down fertile muddy sediments and modifying the coastline by forming deltas. Other great rivers, such as the Ganges, Indus and Brahmaputra, rise in the high mountain belts where plates are colliding. In these collision zones, plate convergence thrusts mountains into the sky, but the steep mountain slopes and high rainfall encourage weathering and erosion. Consequently, sites of mountain building generate prodigious amounts of sediment, which get washed away to be deposited elsewhere. For example, the Bay of Bengal contains several million cubic kilometres of sediment derived from the Himalayan mountains. The fate of sediments is to be eroded and redeposited, or to be turned into sedimentary rocks. Those rocks can then be caught up in subduction zones, either entering the mantle, being forced onto the edge of a continental plate, or being trapped and metamorphosed in a continental collision.

Although the rock cycle moulds the geology of the crust, it does not operate in isolation. The weathering of rock is influenced by climatic effects. The transport and deposition of sediment varies according to the strength of water currents, wind speed or glacial action. The rock cycle is also greatly influenced by organisms. An obvious example is the formation of limestones from corals and other organisms, which form calcium carbonate skeletons or protective structures. Another way in which organisms affect the rock cycle is through weathering. Rocks are partly broken down by the physical action of wind and frost, and by chemical decomposition by rainwater containing dissolved carbon dioxide, but this is greatly accelerated by plant roots, micro-organisms, and other living things in the soil.

In Activity A, 'Rocks and radioactivity', you will be examining various rock samples and using your observations to deduce how the rocks formed (using ideas contained in Sections 2.3 to 2.5) and also placing them in the context of the processes involved in the rock cycle.

2.9 Changing sea-level

At Residential School, as part of Activity C, 'Investigating the environment', you will see sedimentary rocks that reveal how environmental conditions in Britain's geological past were extremely different from those of the present day (in fact 'Britain', like the rest of the Earth's geography is transitory when viewed in terms of the very long span of geological time). As well as evidence from sedimentary rocks (using ideas along the lines of those introduced in Section 2.4), recent landforms also indicate that in the more recent geological past (within the Quaternary Period), sea-level was not the same as it is at present.

- What deposits could we look for that might suggest that sea-level was once higher than at present?
- Beaches today generally have a mixture of pebbles or sand, along with shells and shell fragments. So, if we found such a deposit perched on a cliff above the present high-water mark, we could infer that since it was deposited, there has been a fall in sea-level, leaving behind a raised beach^G.

Raised beaches, sometimes backed by an ancient cliff line, are common features around Britain's coast. Before explaining this observation, we need to ask ourselves has the sea-level fallen, or has the land risen? Both are feasible. For instance, sea-level can fall relative to a fixed landmass if the volume of the oceans shrinks due to cooling and the formation of land-based ice-caps. Similarly, it will rise if these ice caps melt. It so happens that since 2 to 3 million years ago the Earth has been experiencing an ice age — a prolonged period of alternating glacial and interglacial periods characterized by the presence and absence of large polar ice-caps. We are living in an interglacial period that became established about 10 000 years ago. As the volume of ice increases and decreases throughout an ice age, so the volume of the oceans decreases and increases. Some raised shorelines can therefore date from a previous interglacial period when the ocean volume was larger, and hence sea-level was higher than today's.

Nature is not quite so straightforward, however, because the development and removal of thick ice sheets on land surfaces has an effect on the elevation of the land, relative to sea-level. In the last glacial period much of Britain was covered by an ice sheet up to 1.5 km thick for some 60 000 years. As this ice sheet melted, Britain was relieved of the enormous weight of the ice sheet, allowing the land to 'bounce back' in the manner of a ship rising out of the water when a heavy cargo is unloaded, with the result that the land moved relative to sea-level, producing raised shorelines. A complex interplay of changing climate and Earth movements is therefore involved in producing many of the geologically recent coastal landforms of the British Isles. In addition, we can remark that concerns about contemporary sea-level rise relate to yet a further process — that of rapid ocean warming accompanied by the expansion of seawater volume.

2.10 What to do next

You should now read through Sections 1 and 2 of the workbook for Activity C, 'Investigating the environment', for the locations that you will visit during the Residential School. Section 1 outlines the guidelines for carrying out fieldwork and introduces you to the geology of your field area.

2.11 Summary of Section 2

Rocks are classified into three types according to how they were formed. Igneous rocks are formed by crystallization from the molten state; sedimentary rocks are deposited at the Earth's surface from water, air or ice; and metamorphic rocks are rocks of any origin that have been subsequently transformed (metamorphosed) by heat and/or pressure, often several kilometres below the Earth's surface.

Rocks are generally either crystalline, i.e., formed of interlocking mineral crystals, or fragmental, i.e., formed of mineral or rock fragments compacted and cemented together by later mineral growth. Most igneous and metamorphic rocks are crystalline

whereas most sedimentary rocks are fragmental. Most metamorphic rocks have a foliation or mineral banding, and this distinguishes them from igneous rocks. The presence of fossils usually indicates a sedimentary rock.

Slowly cooled magmas produce large crystals, whereas rapidly cooled magmas produce small crystals. The grain size of an igneous rock is therefore an indication of cooling rate. Igneous rocks are classified according to grain size and the minerals present.

Metamorphic rocks may be produced around the margins of igneous intrusions (contact metamorphism) or throughout large volumes of rock in mountain building events where continental plates collide (regional metamorphism). Regional metamorphism produces foliated or banded rocks, whose grain size increases with increasing pressure and temperature (depth).

The surface environment in which a given sedimentary rock formed can be deduced from the character of the sedimentary grains and the types of any fossils and sedimentary structures present. For example, a large grain size indicates that the energy of the depositional environment was high. Well-sorted, well-rounded grains indicate deposition from the air (wind transport) whereas water-laid sediments have poorer sorting and less rounded grains.

Exposures of rocks display the relationships between different rock types and the arrangement in space of those rocks. This gives additional information about how the exposed rocks were formed. The subsequent geological history can be deduced by identifying evidence for geological events that have altered earlier rocks (for example, folding will tilt originally horizontal beds, faulting will offset originally continuous beds, and a break in sedimentation will give an unconformity).

Sedimentary, metamorphic and igneous rocks are produced within a continuously operating cycle — the rock cycle — which is the path taken by Earth materials in response to chemical, biological and physical changes acting on rocks. Those parts of the rock cycle that produce new sedimentary rocks involve the weathering and erosion of pre-existing rocks, followed by transport and deposition of eroded mineral grains and rock fragments. Certain other sedimentary rocks are formed by biological processes, such as the accumulation of shells. In both cases these processes occur on the surface of the Earth. Most igneous and metamorphic rocks are produced near lithospheric plate boundaries.

The presence of raised beaches is evidence of geologically recent sea level change. The processes involved are a combination of climatically-influenced growth or melting of ice-caps and uplift of land in response to unloading when ice sheets are removed by melting.

Now that you have completed Section 2 you should be able to:

- explain the difference between a mineral and a rock;
- describe the textural differences between igneous, sedimentary and metamorphic rocks;
- account for these differences in terms of the processes that produce these rocks;
- classify igneous rocks according to their grain size and mineralogical composition;
- recognize the difference between a body fossil and a trace fossil;
- sketch a rock exposure and identify faults, folds and joints;

- suggest a sequence of geological events that can best explain the features observed in a rock exposure;
- relate processes of the rock cycle to a plate tectonic setting;
- describe the causes of sea-level changes and evidence for these changes;
- understand how to use a hand lens.

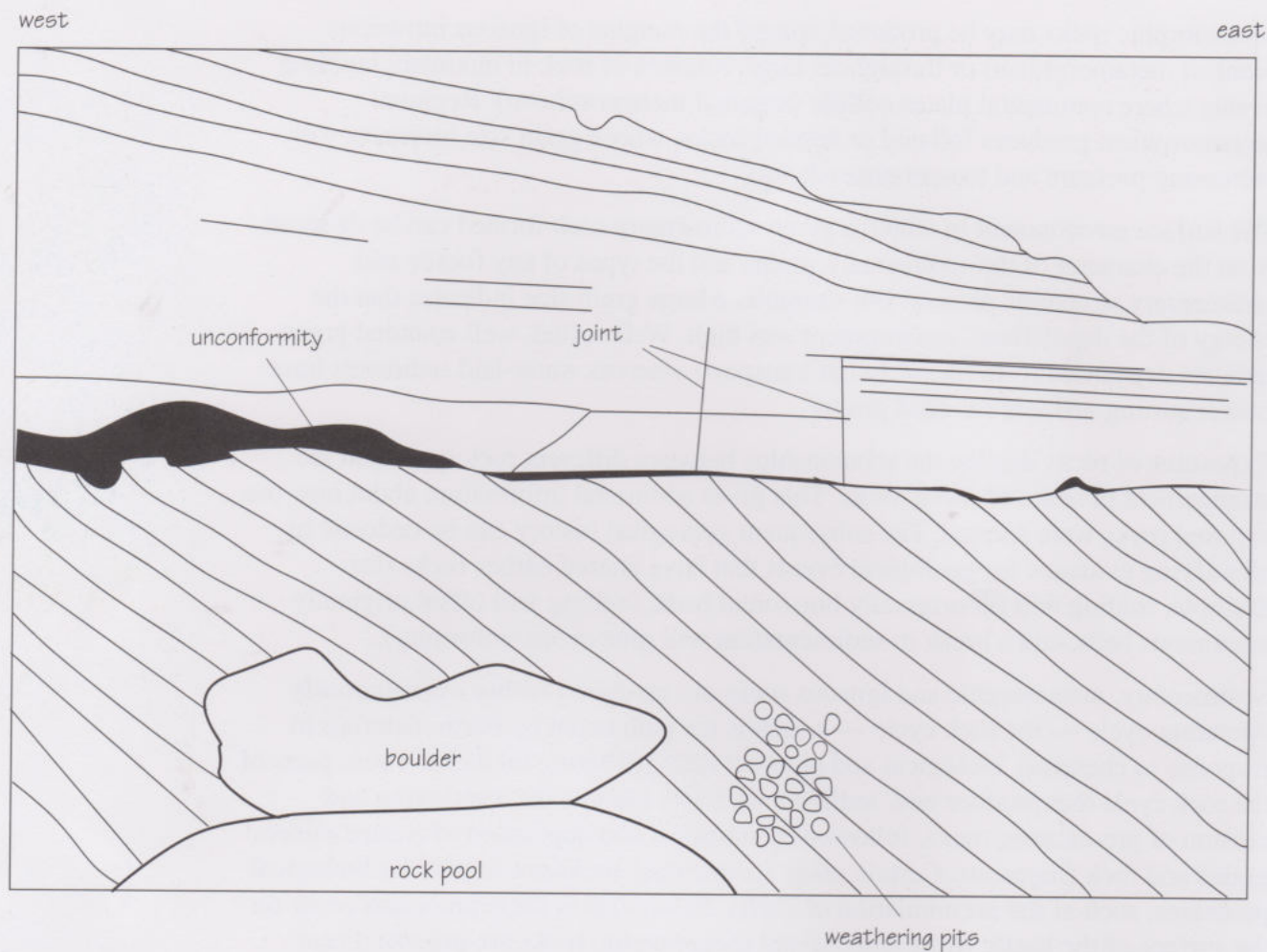


Figure 2.20 A sketch of the exposure shown in Figure 2.16a. The exposure is about 10 m high.

The physics of atoms, energy and radiation

3

3.1 Introduction

One of the fundamental questions that we can ask about the world around us is ‘what is the nature of matter?’ The notion that matter is made up from building blocks called atoms^G is an idea going back to ancient Greece, although scientific evidence for the existence of atoms is a much more recent development. Furthermore, it has only been in the last hundred years or so that scientists have started to understand the properties of atoms in any detail.

Two of the experimental activities that you will encounter at Residential School depend on the properties of atoms. In Activity B, ‘Analysing our environment’, you will see that the electron structure^G of atoms allows scientists to identify which types of atom might be present in a sample of material. In Activity A, ‘Rocks and radioactivity: energy in the Earth’, you will find that an important source of energy within the Earth is the radioactive decay of nuclei of individual atoms.

3.2 The building blocks of the world

An atom is a very small structure, about 10^{-10} m across (see *SGSG* page 359 if you are not sure of the meaning of numbers written as powers of ten, such as 10^{-10} .) This very small size is the main reason why it was historically so difficult to obtain direct evidence for the existence of atoms. However, with the benefit of modern techniques in microscopy, it is possible to form images of individual atoms (see Figure 3.1). There are only about a hundred or so different types of atom, and these types are termed the chemical elements^G, examples being hydrogen, oxygen, carbon and iron.

3.2.1 Atoms and electron structure

At one time it was believed that atoms were fundamental particles, i.e. that they had no internal structure. It is now recognized that the atom is itself made from smaller particles; the electron^G, proton^G and neutron^G. Experiments carried out at the start of the 20th century revealed that most of the mass of the atom is concentrated in a central nucleus^G, which we now know to be made up of protons and neutrons. The nucleus is much smaller than the atom as a whole, typically being about 10^{-14} m across. We shall leave discussion of the nucleus and its interactions until Section 3.4.

The other component of the atom is a distribution of electrons that surrounds the nucleus. This structure is shown schematically in Figure 3.2.

Electrons are far less massive than protons or neutrons, and so they make only a very small contribution to the mass of the atom. However, it is the behaviour of the electrons in an atom that determines many of its characteristic properties. Indeed, the basis on which the science of chemistry is built is the idea that atoms combine together in a way that depends entirely upon their electronic structures.

One of the properties of the atom that you will investigate in Activity B is its ability to emit light of very specific colours. As you will see in the next section, this phenomenon reveals much about the inner workings of atoms as well as being a very useful tool for identifying them.



Figure 3.1 A surface map obtained with a scanning tunnelling microscope. Each blob represents an atom.

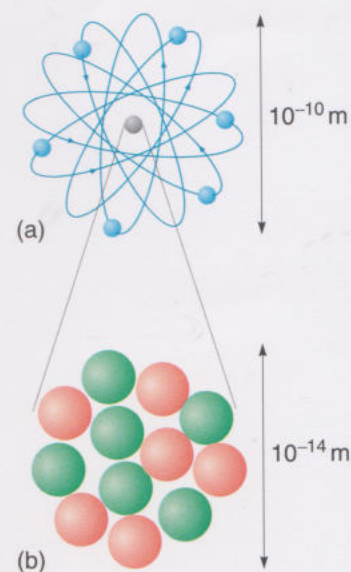


Figure 3.2 The constituents of a typical atom: (a) electrons move around the tiny nucleus, which is the core of the atom; (b) the nucleus is made up of particles called protons and neutrons.

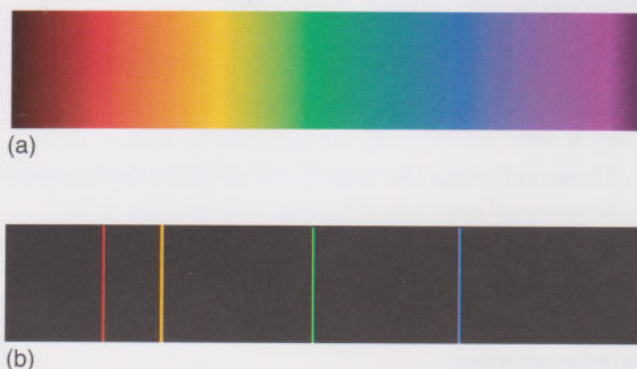
Question 3.1 What fraction of the typical diameter of an atom is the diameter of its nucleus? (Use the information given in Sections 3.2 and 3.2.1, and express your answer in the form $1/(\text{a number})$. If you are not sure how to multiply and divide numbers like 10^{-10} and 10^{-14} , see *SGSG* page 363.) ◀

3.3 Spectral lines: an atom's signature

3.3.1 Spectra and spectral lines

When a beam of white light from the Sun is passed through a prism it is broken up, or dispersed, into a range of colours that form a pattern similar to a miniature rainbow. Such a band of colours is referred to as a spectrum^G (plural spectra). The spectrum of white light from a conventional light bulb comprises an uninterrupted band of colours known as a continuous spectrum^G, an example of which is shown in Figure 3.3a. We can use a similar technique to study the light emitted by specific types of atom. An example of the type of source that could produce light in this way is a sodium street light. In this case we see a spectrum that is very different from the continuous spectrum, as it consists of very narrow lines of specific colours (see Figure 3.3b). This spectrum is termed an emission spectrum^G and the lines in it are called spectral lines^G. Sodium has very prominent yellow spectral lines as well as fainter green, red and blue lines.

Figure 3.3 (a) A continuous spectrum from a beam of white light. (b) The emission spectrum of sodium.



How does this characteristic pattern of lines of different colours relate to the processes that take place within individual atoms? A starting point is to note that light carries energy from one place to another. A simple example of this is the ‘solar-powered’ calculator, in which the energy carried by light is converted into electrical energy. What is less obvious is that light carries energy in tiny ‘packets’. These ‘packets’ are called quanta (singular quantum^G). In fact, there is a simple connection between the energy of these quanta and the colour of light — quanta of red light have lowest energy and quanta of violet light have the highest energy. When individual atoms produce light, they produce individual quanta. These ‘packets’ of light are often referred to as photons^G.

One of the cornerstones of physics is the principle that energy is conserved. If, in an experiment to observe light from a particular type of atom we notice that photons of a certain energy (E_{ph}) are produced, then this means that the atom itself must have lost exactly this amount of energy. In other words, there has been a change in the energy of the atom that is equal to the energy of the emitted photon. We can express this mathematically as

$$\Delta E_{\text{atom}} = E_{\text{ph}} \quad (3.1)$$

where ΔE_{atom} is the change in the energy of the atom, and E_{ph} is the energy of the photon. (The Greek letter Δ (delta) is often used in science to indicate a change in a quantity.)

The fact that atoms produce spectral lines of particular energies suggests something about the energy that atoms themselves can have, namely that they can have only specific energies. One way to visualize this is by means of a diagram that shows the energy levels^G that an atom can have. Figure 3.4 shows such a diagram for a hypothetical atom that has just two energy levels: a low energy level (E_{low}) and a high energy level (E_{high}). When the atom loses energy in the form of a photon, it makes a transition from the E_{high} to the E_{low} level. The energy of the photon is the difference between the two energy levels, i.e. $E_{\text{ph}} = E_{\text{high}} - E_{\text{low}}$. The opposite process can also occur. An atom that is in the E_{low} level can absorb a photon and reach the E_{high} level. However only a photon with exactly the right amount of energy ($E_{\text{high}} - E_{\text{low}}$) can cause this transition.

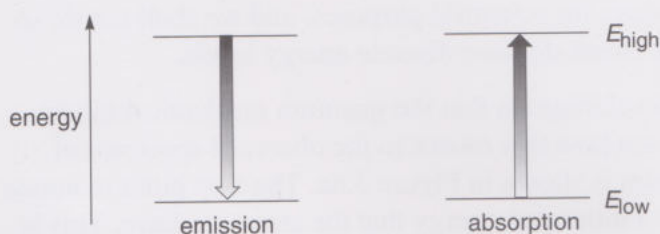


Figure 3.4 Energy level changes for (left) an emission spectral line, and (right) an absorption spectral line.

This two-state atom is much simpler than any real atom, but the important feature is the fact that this atom, like all real atoms, has discrete energy levels. Furthermore, each chemical element has its own distinctive pattern of energy levels. The energies of photons that are seen in the emission spectra of elements are determined by the spacings of these levels, and so the spectrum can be considered to be a unique signature that can be used to identify that particular element.

- A hypothetical atom has three energy levels E_1 , E_2 and E_3 , as shown in Figure 3.5. How many different transitions are possible that would result in the emission of photons? What would be the energies of these photons?
- There are three possible transitions of this atom that will result in the emission of photons. These transitions are: $E_3 \rightarrow E_1$, $E_3 \rightarrow E_2$ and $E_2 \rightarrow E_1$, and result in photons of energies $(E_3 - E_1)$, $(E_3 - E_2)$ and $(E_2 - E_1)$, respectively.

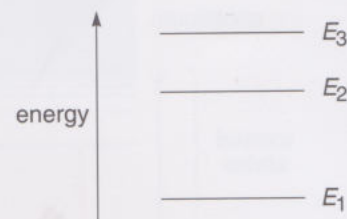


Figure 3.5 An energy level diagram with three levels.

3.3.2 Discrete energy levels and spectra

We have seen that changes between discrete energy levels in an atom give rise to photons with particular energies, and that these photons form a characteristic spectrum. Spectra were first studied in the 19th century, but at that time the origin of spectral lines was a mystery. In the early part of the 20th century, considerable effort was made by physicists to develop a model of the hydrogen atom, because this is the simplest possible atom, having a single proton as the nucleus and a single electron. (The hydrogen atom is unique in that it does not contain any neutrons.) A key observation in understanding the interactions that take place within an atom is the fact that both electrons and protons are electrically charged particles. All protons have a positive charge of the same value (which is called e). All electrons have a negative charge of $-e$. Two bodies that have opposite charges will attract each other by an electric force. So a model of the hydrogen atom can be developed in which a

negatively charged electron is bound by this electric force to the positively charged proton. On this basis we can get a feel for how such a system might respond to changes in energy. If we imagine a hypothetical atom in which the electron is initially close to the proton then, to transform this atom into a state in which the electron is more distant from the proton, it is necessary to do work to increase the separation of these two particles. So in order to increase the separation of the electron from the proton, it is necessary to supply energy to the atom. The converse is also true: when the separation of the electron and the proton decreases, energy is lost from the atom. This view of the atom explains how the atom can have different energies, but it offers no explanation for the fact that the atom can have only a discrete set of energy levels.

As a result of trying to solve this and other problems, a very successful, yet revolutionary theory was developed in the first quarter of the 20th century. This theory is called quantum mechanics, and it is remarkable because it requires us to discard many of our 'common-sense' notions about the way in which particles behave. It is not necessary to delve into the quantum origins of the discrete energy levels in atoms in order to use spectra for scientific purposes, and we shall not do so here. We shall simply accept that atoms do have discrete energy levels.

It is useful to look at the energy level diagram that the quantum mechanical model of the hydrogen atom predicts to see how this relates to the observed spectrum of hydrogen. The energy level diagram is shown in Figure 3.6a. The first point to notice about this diagram is that there is a minimum energy that the atom can have. This is called the ground state^G of the atom and it corresponds to a state where the average distance between the electron and the proton has the smallest value of any energy level. This energy level is given the label ' $n = 1$ ', where the number n is called the principal quantum number^G. At higher energies than the ground state there are excited states^G, labelled $n = 2, 3, 4$ and so on, which correspond to increases in the

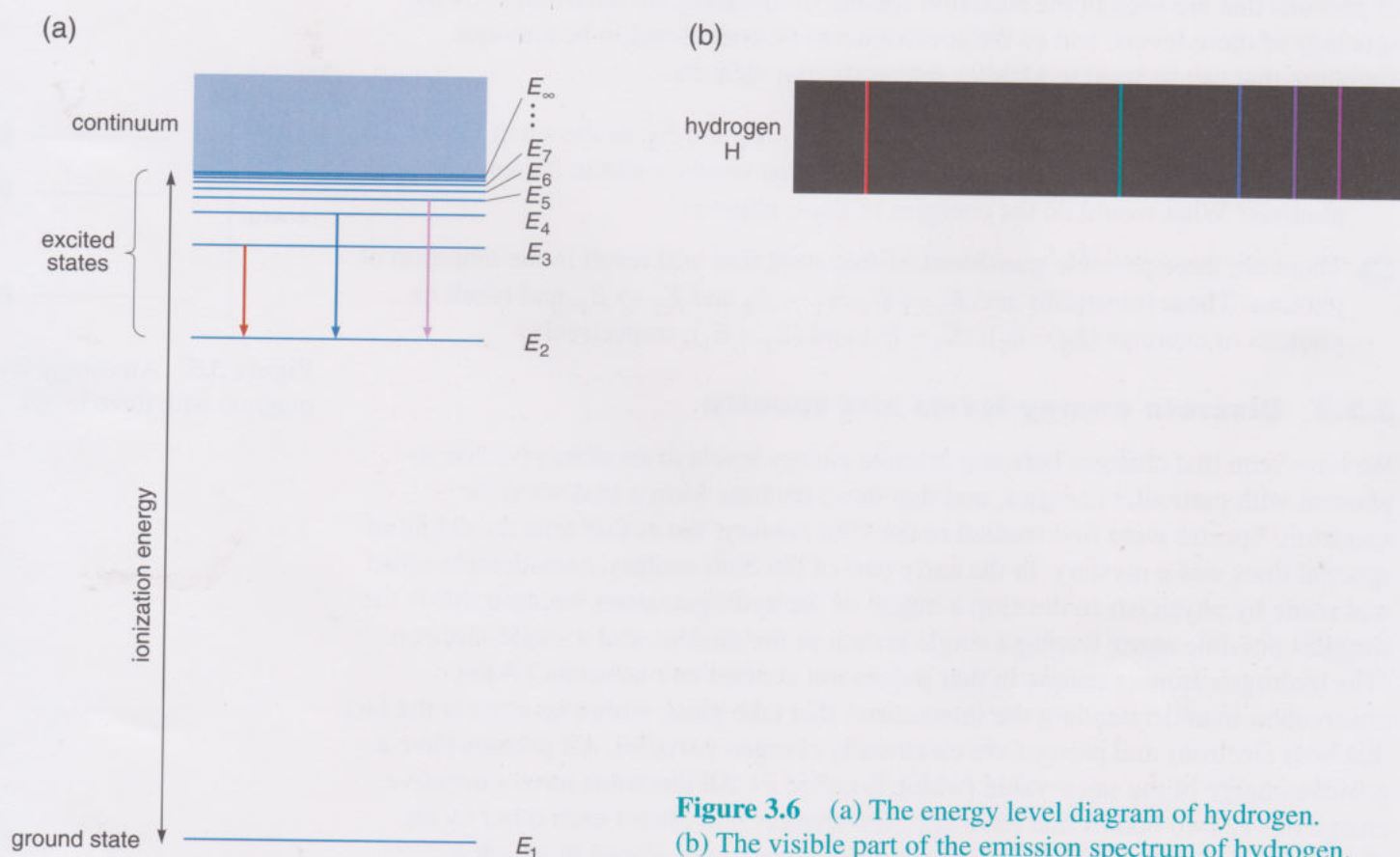


Figure 3.6 (a) The energy level diagram of hydrogen. (b) The visible part of the emission spectrum of hydrogen.

average separation of the electron and proton. Note that the energy levels are not evenly separated: in fact, as n increases, the energy difference between successive levels gets progressively smaller, until we reach an energy at which the electron is free from the influence of the proton. At this energy the electron is no longer a part of the atom. The process of removing an electron from an atom is called ionization^G, and an atom that has lost one (or more) electrons so that it is no longer electrically neutral is called an ion^G.

An atom emits a photon when the atom makes a transition from one energy state to a lower energy state. For instance, if the hydrogen atom changes from the $n = 3$ to the $n = 2$ state, then a photon is emitted that corresponds to a red line in the spectrum. The transition from the $n = 4$ state to the $n = 2$ state results in a photon of higher energy, because the energy difference between the $n = 4$ level and the $n = 2$ level is greater than the energy difference between the $n = 3$ level and the $n = 2$ level. The transition from the $n = 4$ state to the $n = 2$ state results in a spectral line that is in the blue part of the spectrum. For the transition from the $n = 5$ state to the $n = 2$ state, there is a photon of yet higher energy emitted, in this case corresponding to a violet spectral line. This trend continues, and as Figure 3.6b shows, there is a distinctive pattern of lines at particular energies that correspond to transitions to the $n = 2$ state.

You might be wondering about where in the spectrum a line from a transition to the $n = 1$ level might appear. The lowest energy photon from such a transition arises from a change from the $n = 2$ state to the $n = 1$ state. This corresponds to a photon that is more energetic than a photon of violet light, and so it lies in the ultraviolet part of the spectrum. The human eye is unable to see this part of the spectrum. Similarly, transitions between closely separated levels give rise to photons with less energy than photons of red light, and these photons lie in the infrared part of the spectrum, so again we are unable to see them.

In summary, an atom that is above its ground state produces photons of energies that reflect the separation between energy levels in that atom. Here we have concentrated on the spectrum produced by hydrogen atoms. Other chemical elements have energy level diagrams that are different from that of hydrogen, and there is a unique and distinguishing spectrum for each element.

3.3.3 The production of emission spectra

In order to get an atom to produce photons, the atom must be raised above its ground state, and this is done by supplying energy to the atom. One way of doing this is the opposite of the emission process that we have just described. We have seen that when an atom loses energy it does so by emitting photons with energies that correspond to the differences between energy levels. In the opposite process, an atom in one energy state can transform to another higher energy state by absorbing a photon. However, the incoming photon must have an energy that corresponds to the difference between the two energy levels for this to occur. Another method of raising energy states is to collide atoms together. In this case, some of the energy of the collision may be absorbed by one of the atoms, causing the atom to change to a higher energy state. In practice, this can be achieved readily by heating a gaseous sample containing the atoms that one wishes to study. This can be done by introducing a sample into a flame. In Activity B, 'Analysing our environment', gas burners are used to produce flames that contain samples of metal ions that it will be your task to identify.

3.3.4 Spectroscopy: measuring spectral signatures

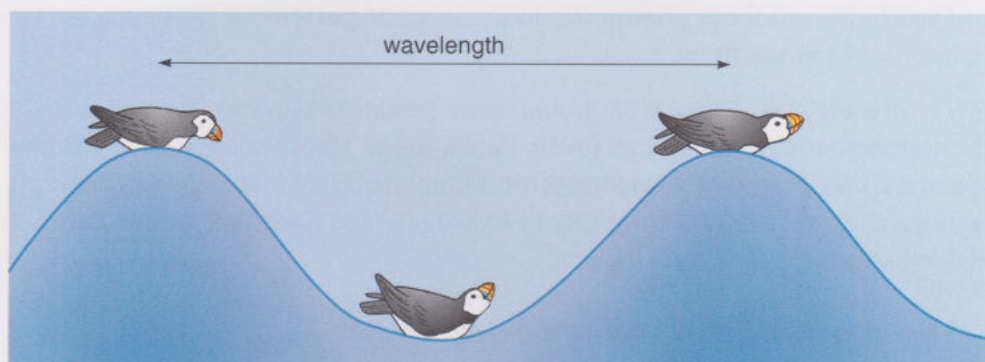
The technique of identifying which metal atoms are present in a sample of material by analysis of the spectrum of emitted photons is called spectroscopy^G. Scientists often want to discover which elements may be present in some sample of material, and spectroscopy provides the means to this end in a wide range of applications, such as the analysis of drinking water and investigation of the composition of stars. In Activity B you will identify elements from their spectra. The following section describes some of the theory behind the technique of spectroscopy that you need to know before starting the practical work.

3.3.5 Energies and wavelengths

You have already seen that elements have a spectral signature that consists of spectral lines at characteristic energies. In practice it is rather difficult to measure the energies of photons directly, but there is an indirect way of measuring photon energy which relies on that fact that light has wave-like behaviour.

If you have ever spent any time at the seaside you will probably be aware of the most important features of waves. A wave^G is a periodic or repeating disturbance that carries energy from one place to another. A typical profile of the disturbance caused by a wave on water is shown in Figure 3.7. One of the characteristics of the wave is the distance between two consecutive peaks (or troughs). This is called the wavelength^G and is usually denoted by λ (the Greek letter lambda). In the case of light the wavelength is a very short distance, only about 4×10^{-7} m or so, and this is one reason why the wave-like nature of light is not obvious to us.

Figure 3.7 An illustration of wavelength.



The exact nature of light need not concern us too much here, but it is useful to know that light is one example of a phenomenon known as electromagnetic radiation^G. The wavelengths of electromagnetic waves cover an enormous range: from over 10^4 m for radio waves to less than 10^{-16} m for gamma rays^G. You will meet gamma rays again in Section 3.5 when energy changes in the nucleus of an atom are discussed.

The wavelength of all electromagnetic radiation has a simple relationship with the energy carried by its photons. Mathematically this relationship can be written

$$E_{\text{ph}} = \frac{hc}{\lambda} \quad (3.2)$$

where h is a quantity known as the Planck constant (named after the German physicist Max Planck) and c is the speed of light in a vacuum.

- Both h and c are constants. According to Equation 3.2, what determines the energy of a photon?

- The energy of a photon is determined by the wavelength of the light.

It can also be seen from Equation 3.2 that if the wavelength is very large then the photon energy is very small and vice versa. We have seen that atoms produce spectral lines of given energies, but we could equally well say that the atom produces spectral lines at particular wavelengths.

- Yellow light comprises photons of lower energy than the photons that comprise green light. Which colour has the longer wavelength?
- From Equation 3.2 it can be seen that energy decreases as wavelength increases. So if yellow light has lower energy photons than green light, then yellow light must have a longer wavelength than green light.

3.3.6 Measuring the wavelength of spectral lines

In order to measure the wavelengths of spectral lines, it is necessary to generate spectra in such a way that we can relate the position of a line in the spectrum to its wavelength. It turns out that using a prism to form the spectrum is not very well suited to this task. There is however a different type of optical device, called a diffraction grating^G, which produces a spectrum in which it is straightforward to relate position in the observed spectrum to wavelength. A diffraction grating is simply a series of very narrow, evenly spaced slits. To form the spectrum, the grating is set up so that the light that is to be analysed hits the grating at an angle of 90° , as indicated in Figure 3.8.

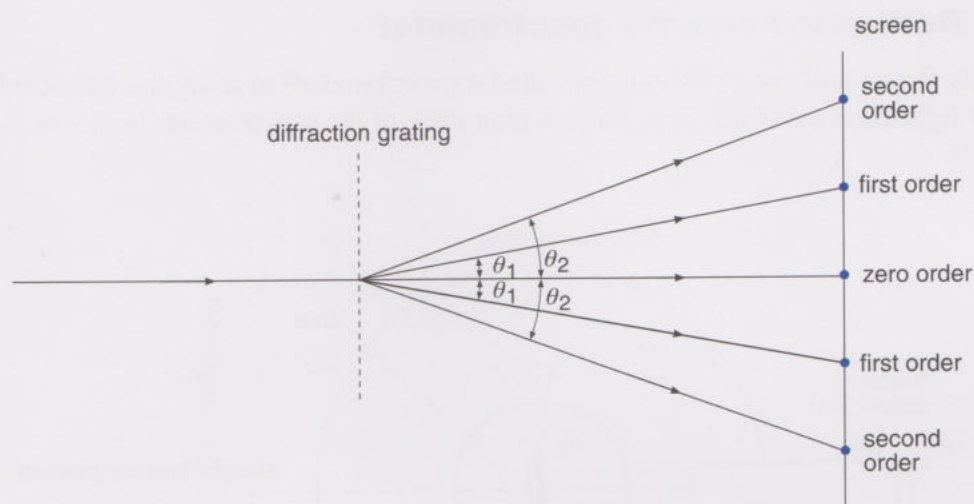


Figure 3.8 The pattern formed by a diffraction grating.

The grating acts in such a way that light of a given wavelength forms more than one beam after it has passed through the grating, forming what is called a diffraction pattern. Let us consider what happens if we shine light of a single wavelength (or monochromatic^G light) onto the diffraction grating. Some light passes straight through the grating; this is called the zero order beam. Another beam is produced at an angle θ_1 from the straight through position. This is called first order diffraction, and an identical beam is produced at an angle θ_1 in the other direction from the straight through position. The importance of this diffracted beam is that the angle that

the beam emerges at (called the angle of diffraction^G) is related to the wavelength, λ , of the light by the equation

$$\sin \theta_1 = \frac{\lambda}{d} \quad (3.3)$$

where d is the grating spacing^G, the distance between the centres of two adjacent slits on the grating. This means that if we know the value of d for our grating, then assuming that we can measure the angle of diffraction, we can calculate the wavelength of the light (see Appendix 1 if you would like a reminder about the use of sines).

The diffraction pattern includes other beams of light of this wavelength. At a higher angle still from the straight through position there will be a second diffracted beam, called the second order diffraction. In this case too, there is a simple mathematical relationship between the angle of diffraction, which we now call θ_n and the wavelength:

$$\sin \theta_n = \frac{n\lambda}{d} \quad (3.4)$$

The quantity n is called the order of diffraction^G. For first order diffraction $n = 1$ and for second order diffraction $n = 2$. It is possible to generate third ($n = 3$) and higher orders of diffraction, but you will not do this in Activity B.

Question 3.2 A beam of light is diffracted by a grating that has a spacing of 1.667×10^{-6} m. The spectrum produced by the grating has one first order diffraction beam at an angle of 19.8° and another first order diffraction beam at an angle of 23.4° . Calculate the wavelengths of both of the components of the light.

3.3.7 Getting to know the spectrometer

In Activity B you will use an instrument called a spectrometer^G to study the diffracted beams of light from a diffraction grating. A plan view of the spectrometer is shown in Figure 3.9.

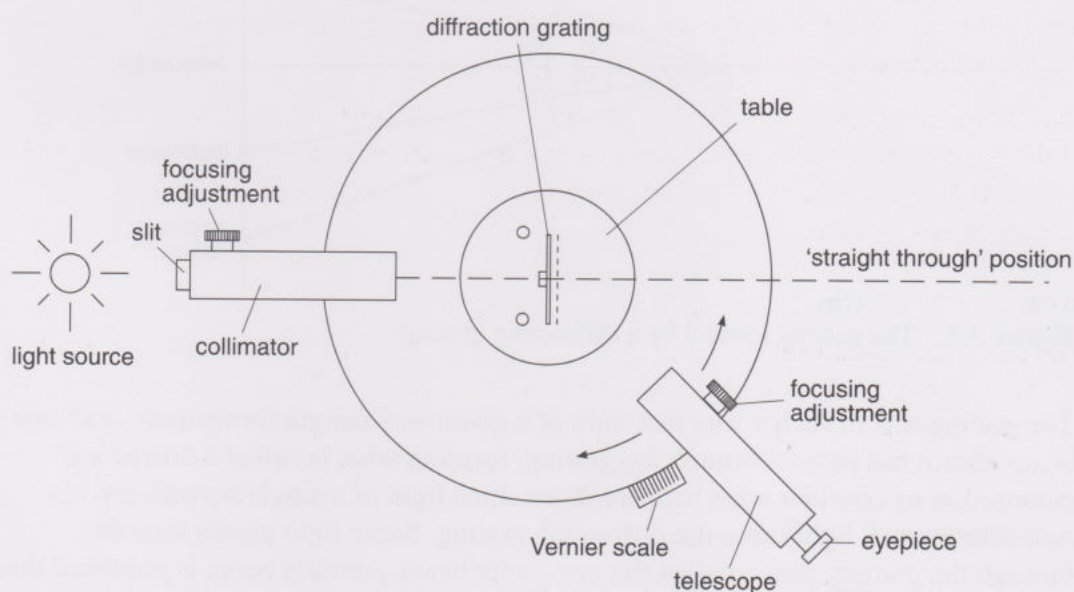


Figure 3.9 A schematic illustration of a spectrometer.

The spectrometer itself consists of a table on which the diffraction grating may be mounted. The light to be analysed passes through an instrument called the collimator^G before striking the diffraction grating at an angle of 90° .

Before striking the grating, the beam of light should not be spreading out or getting narrower; this is called a parallel beam. The collimator contains a series of lenses which, when properly focused, will produce a parallel beam. In addition to being a parallel beam, it is a great advantage if the light to be analysed is in a narrow beam. This is because the diffracted beams are the same width as the incoming beam, and the narrower the diffracted beam the more precisely the angle of diffraction can be measured.

After passing through the grating, light is diffracted into beams that emerge at angles that depend on the wavelength of the line. In order to see these emergent diffracted beams there is a telescope^G, which can be rotated around the table. An angular scale that runs around the table allows the position of the telescope to be determined for any observed beam.

3.3.8 An outline of the parts of Activity B relating to spectroscopy

A full description of Activity B is given in the workbook, but the main stages that deal with spectroscopy are as follows. The overall aim is to identify metals by investigation of the visible light that they emit when heated to high temperatures.

The first part of the activity is an investigation by eye of the colours of flames of unknown samples. This analysis is then taken further by observing the spectra of these flames using a handheld spectroscope^G (a device that produces a spectrum, but which cannot easily be used to measure the wavelength of any spectral lines). The activity then moves on to using a spectrometer of the type described in Section 3.3.7. Before the spectrometer can be used, the collimator and telescope need to be properly focused, and the spectrometer needs to be calibrated^G. The aim of the calibration process is to determine the value of the grating spacing d of the particular grating that you will be supplied with. This involves using a spectral line as a wavelength standard, and for this activity, we use a bright yellow spectral line of sodium. Finally, the calibrated spectrometer is used to measure the wavelengths of spectral lines obtained from flames that contain unknown metals. These metals can then be identified from the pattern of observed spectral lines.

You should now read about the investigation in more detail. It is described in Part B of the Activity B workbook. To conclude the introduction to spectroscopy, you should carry out Activity 3.1 in which you will analyse spectrometer data with the aim of calibrating an instrument.

Activity 3.1 Calibrating a diffraction grating

The aim of this data analysis activity is to find a value for the grating spacing, d , of a diffraction grating, from some data which we will give you. The activity should also give you confidence in plotting and using graphs. The activity is broken down into several distinct tasks. If you have any doubts about the way in which you have answered the questions related to the tasks, you should compare your answers with those given at the end of this book.

Prior to its use to measure the wavelength of spectral lines, the spectrometer is calibrated by using a monochromatic source that produces light at a known fixed wavelength of 632.8 nm. (The nanometre^G (nm) is equal to 10^{-9} m, and is a unit used frequently when discussing the wavelength of light.) Diffracted beams are observed up to and including the fourth order. The angles of diffraction are shown in Table 3.1.

Table 3.1 Some angles of diffraction, for use in Task 1.

n	$\theta_n/^\circ$	$\sin \theta_n$
0	0.0	
1	12.1	
2	25.1	
3	39.2	
4	57.5	

Task 1

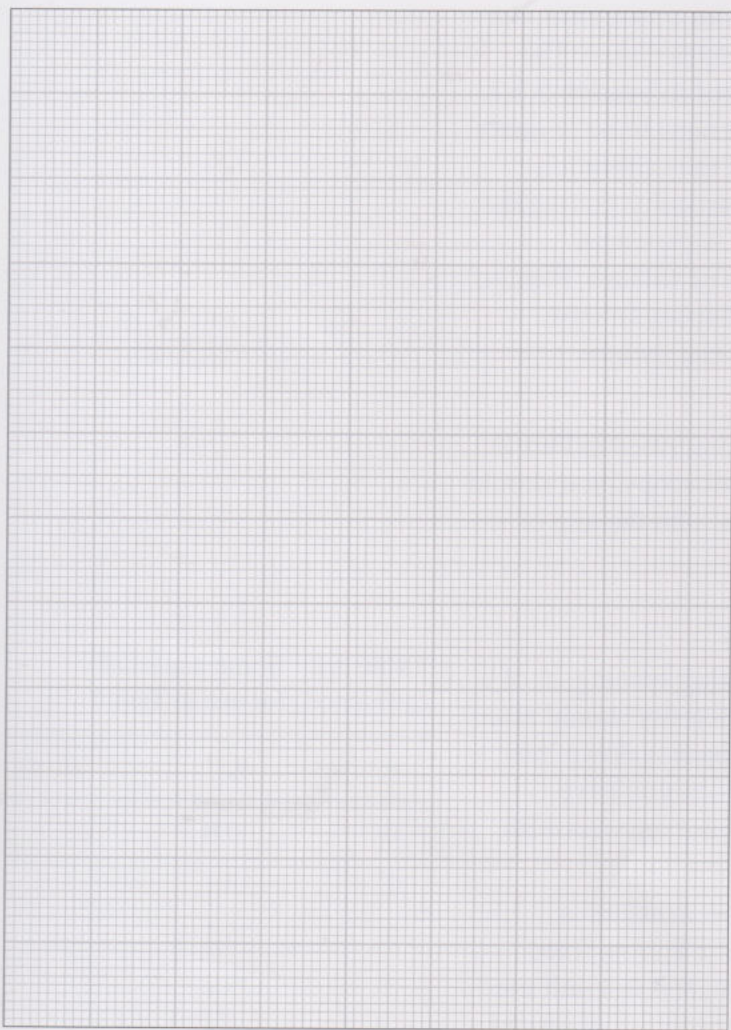
We start from the diffraction equation

$$\sin \theta_n = \frac{n\lambda}{d} \quad (3.4)$$

- (a) Which of the terms in Equation 3.4 are: (i) known, (ii) unknown, or (iii) measured in the experiment?
- (b) Does Equation 3.4 describe a straight line relationship between the two measured quantities in the experiment? If not, how can a straight line relationship be obtained? (Hint: remember that a straight line graph going through the origin can be described by an equation of the form $y = kx$, where x and y are the variables plotted in the graph, and k is a constant. Appendix 2 gives more detail about the equations of straight line graphs.)
- (c) Suppose that we plot a graph of $\sin \theta_n$ against n . How would we find the value of d , the grating spacing, from such a graph?

Task 2

We are now going to calculate the quantities that are needed to plot the required graph. Complete the blank column of Table 3.1 by calculating the sines of the angles of diffraction.

**Task 3**

We now want to plot the straight line, measure its gradient and calculate d .

- (a) On the piece of graph paper provided in Figure 3.10, plot a graph of $\sin \theta_n$ (vertical axis) against n (horizontal axis). Draw a straight line that is the best fit to the data on this graph. (See *SGSG* page 376 for some hints on graph plotting.)
- (b) Now find the gradient of the best fit line. (See *SGSG* page 384 for guidance on calculating gradient.)
- (c) Calculate a value for d . Express your answer in nanometres and in metres. ◀

Figure 3.10 Graph paper for use in Activity 3.1

3.4 Into the nucleus

We have already discussed the fact that atomic nuclei contain protons and neutrons. Each proton carries a positive charge exactly equal in size to the negative charge carried by an electron and, as the name suggests, neutrons are electrically neutral. This implies that each electrically neutral atom must contain the same number of protons as electrons. It is this number, called the atomic number^G, Z , which determines an element's chemical properties — if the number of protons in an atom is altered, it becomes an atom of a different element. Note, for example, that hydrogen atoms always contain one proton, copper atoms always contain 29 protons, and gold atoms always contain 79 protons.

The neutrons inside a nucleus do not contribute to its charge, but they do contribute to its mass. Some elements exist in several different forms, known as isotopes^G, each with the same atomic number (i.e. the same number of protons), but with different numbers of neutrons and hence different masses. For example, there are three naturally occurring isotopes of uranium; all three have 92 protons (it is this number that identifies the element as uranium), but one has 142 neutrons, one has 143 neutrons and one has 146 neutrons. The total number of neutrons and protons (known collectively as nucleons^G) in a nucleus defines its mass number^G, A . Thus the mass numbers of the three naturally occurring isotopes of uranium are 234 (92 protons plus 142 neutrons), 235 (92 protons and 143 neutrons) and 238 (92 protons and 146 neutrons) respectively. The isotopes are commonly referred to as uranium-234, uranium-235 and uranium-238.

Sometimes it is useful to represent atoms using symbols that show both the mass number and the atomic number, and hence the number of protons and neutrons in the nucleus. Against the symbol for the element, we write the atomic number as a preceding subscript, and the mass number as a preceding superscript. The helium atom has atomic number 2 and mass number 4, so we write it as ${}^4_2\text{He}$.

● How many protons and neutrons are there in the helium atom?

● The atomic number is 2, so there are two protons.

The mass number gives the number of protons plus neutrons, so the number of neutrons is the mass number minus the atomic number = $4 - 2 = 2$.

Figure 3.11 gives the symbolic representation for the three isotopes of uranium which we considered earlier.

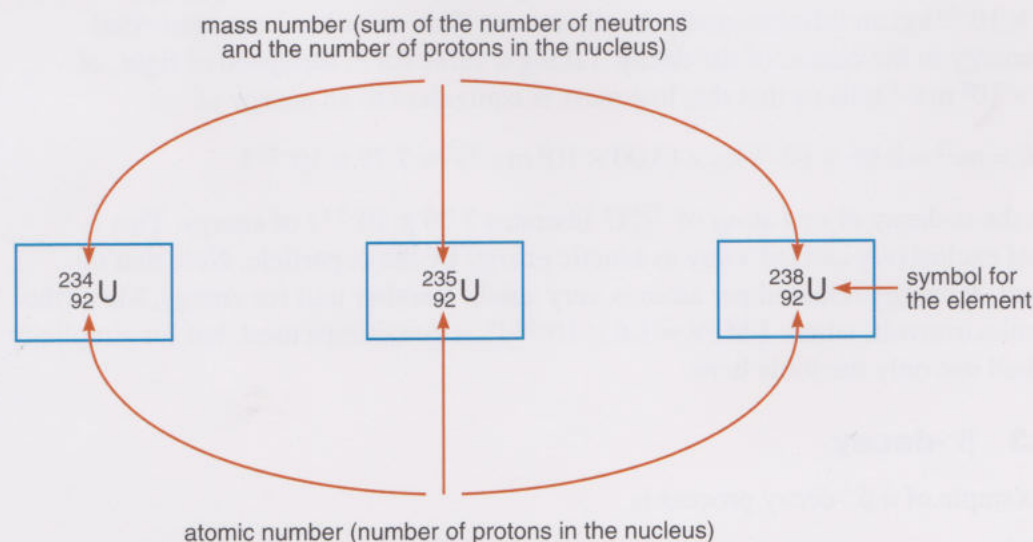


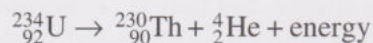
Figure 3.11 Three isotopes of uranium, in symbolic representation. Note that the atomic number (and hence the number of protons) is the same for each, indicating that they are the same element; but the mass number is different, indicating that they are different isotopes.

3.5 Radioactive decay

In Activity A, 'Rocks and radioactivity', you will investigate radioactive decay^G, the process whereby certain unstable atomic nuclei break up or decay spontaneously, with the release of energy. There are various types of decay possible, but the three that are most important in this Activity are those known as alpha decay^G (α -decay), beta-minus decay^G (β^- -decay) and gamma-decay^G (γ -decay).

3.5.1 α -decay

An example of an α -decay process is



In words, a uranium (U) nucleus (the parent nucleus) decays to produce a thorium (Th) nucleus (the daughter nucleus), a helium nucleus and energy. The helium nucleus is more commonly known as an α -particle^G. Notice that the number of protons in the product is not the same as that in the parent nucleus, so the product is no longer uranium, but an isotope of thorium. Notice too that the overall mass number and the total charge are the same on both sides of the equation: the mass number on the left (234) is equal to the sum of the mass numbers on the right (230 + 4); the charge on the nucleus on the left (from 92 protons in the uranium nucleus) is equal to the total charge on the nuclei on the right (from 90 protons in the thorium nucleus and two protons in the α -particle). Equations that represent nuclear decays must always balance in this way, i.e.

- Electric charge is always conserved: the net charge on the products of a nuclear decay is the same as the net charge of the original nucleus.
- The mass number is conserved: the total number of nucleons in the products is the same as that in the original nucleus.

3.5.2 Where does the energy come from?

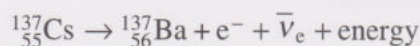
You should be familiar with the idea that energy is conserved in all physical processes. So where does the 'energy' on the right hand side of the equation come from? The answer lies in Einstein's famous equation, $E = mc^2$. Accurate measurements of the masses of the α -particle and the thorium nucleus reveal that their sum is *less* than the mass of the original uranium nucleus by about 8.66×10^{-30} kg, an infinitesimally small amount. This mass has been converted into energy in the course of the decay. Taking a value for c , the speed of light, of 3.00×10^8 m s⁻¹ tells us that this lost mass is equivalent to an energy of

$$E = mc^2 = 8.66 \times 10^{-30} \text{ kg} \times (3.00 \times 10^8 \text{ m s}^{-1})^2 = 7.79 \times 10^{-13} \text{ J}$$

Thus the α -decay of one atom of ${}_{92}^{234}\text{U}$ liberates 7.79×10^{-13} J of energy. This is almost exclusively carried away as kinetic energy by the α -particle. Note that the amount of energy released per atom is very small. Another unit for energy, MeV (the megaelectronvolt, where $1 \text{ MeV} = 1.6 \times 10^{-13} \text{ J}$), is sometimes used, but for simplicity we shall use only the joule here.

3.5.3 β^- -decay

An example of a β^- -decay process is

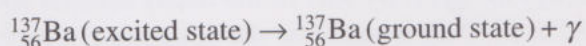


In words, a caesium nucleus decays to produce a barium nucleus, a high energy electron, another particle called an electron antineutrino^G ($\bar{\nu}_e$) and energy. The antineutrino has no charge, negligible mass and is very difficult to detect. The electron given off in the process was not initially recognized as an electron, so it became known as a β -particle^G.

Let's start by checking that the equation balances. The electron carries a negative charge exactly equal in size to the positive charge carried by a proton, and the antineutrino has no charge. Thus the charge on the left-hand side (from the 55 protons in the caesium nucleus) is equal to the charge on the right hand side (from the 56 protons in the barium nucleus and the single negatively charged electron). Also, the left-hand side and the right-hand side both have a mass number of 137. But where has the electron come from? We're talking about a *nuclear* decay, so it has not come from the electrons surrounding the nucleus in the atom. Neither was it in the nucleus in the first place — nuclei don't contain electrons! The correct explanation involves noting that, since the barium nucleus contains one more proton than the caesium nucleus, but the same number of nucleons altogether, it must contain one less neutron than the original caesium nucleus. What has happened is that one of the neutrons in the caesium nucleus has been transformed into a proton, with the emission of an electron.

3.5.4 γ -decay

In contrast to the processes of α - and β^- -decay, γ -decay involves no change in the numbers of neutrons and protons. Instead it occurs in a process which is analogous to the one which gives rise to *atomic* spectra, as discussed in Section 3.3.2 and illustrated in Figure 3.6a. γ -decay occurs when a *nucleus* makes a transition from a state of high energy to one of lower energy, and the transition is accompanied by the emission of a photon, for example:



The photon emitted in the nuclear process has an energy around a million times larger than in the atomic one, so it is a γ -ray photon rather than a photon of visible light. Excited states of nuclei may be created as a result of α -decay or β^- -decay, so γ -decay often accompanies other types of radioactive decay.

Activity 3.2 The decay of uranium

Let's look in more detail at the decay of ${}^{238}_{92}\text{U}$. Over 99% of the uranium in the Earth is present as this radioactive isotope, and we show a partially completed decay scheme for it in Table 3.2. As you can see, ${}^{238}_{92}\text{U}$ doesn't decay straight to a stable product, but does so via a variety of decay processes, each of which gives an intermediate radioactive isotope, until it eventually decays to give a stable isotope of lead. Many of the daughter nuclides produced in this decay are created in excited states and decay via γ -ray emission to their ground state before a subsequent α - or β^- -decay occurs, but the γ -ray emissions are not shown in Table 3.2.

Task 1

Each of the decay processes omitted from Table 3.2 is either an α -decay or a β^- -decay. Use the rules you have learnt about conservation of charge and conservation of mass number to work out which process leads to each product, and thus complete Table 3.2.

Table 3.2 The radioactive decay scheme for uranium-238.

Step	Parent	Decay	Daughter
1	${}^{238}_{92}\text{U}$	α	${}^{234}_{90}\text{Th}$
2	${}^{234}_{90}\text{Th}$	β^{-}	${}^{234}_{91}\text{Pa}$
3	${}^{234}_{91}\text{Pa}$	β^{-}	${}^{234}_{92}\text{U}$
4	${}^{234}_{92}\text{U}$	α	${}^{230}_{90}\text{Th}$
5	${}^{230}_{90}\text{Th}$		${}^{226}_{88}\text{Ra}$
6	${}^{226}_{88}\text{Ra}$		${}^{222}_{86}\text{Rn}$
7	${}^{222}_{86}\text{Rn}$		${}^{218}_{84}\text{Po}$
8	${}^{218}_{84}\text{Po}$		${}^{214}_{82}\text{Pb}$
9	${}^{214}_{82}\text{Pb}$		${}^{214}_{83}\text{Bi}$
10	${}^{214}_{83}\text{Bi}$		${}^{214}_{84}\text{Po}$
11	${}^{214}_{84}\text{Po}$		${}^{210}_{82}\text{Pb}$
12	${}^{210}_{82}\text{Pb}$		${}^{210}_{83}\text{Bi}$
13	${}^{210}_{83}\text{Bi}$		${}^{210}_{84}\text{Po}$
14	${}^{210}_{84}\text{Po}$		${}^{206}_{82}\text{Pb}$ (stable)

Task 2

Each time a single ${}^{238}_{92}\text{U}$ nucleus decays to ${}^{206}_{82}\text{Pb}$ by way of the decay scheme shown in Table 3.2, a total mass of about $9.2 \times 10^{-29} \text{ kg}$ is lost. Use the equation $E = mc^2$ to calculate the corresponding energy (in joules). ◀

3.5.5 When do nuclei decay?

We have referred to some isotopes as stable and others as unstable; the unstable nuclei are those that decay. But when will this decay happen? The simple answer is that we don't know. Radioactive decay is an intrinsically random process and we can never know when an individual nucleus will decay. However, if we take a large enough number of nuclei, we can use the concept of half-life^G (the time taken for half a radioactive sample to decay) as a measure of how rapidly the isotope is likely to decay. Half-life is a useful and important concept because it gives an indication of the persistence of a radioactive isotope in the environment. For some uses, for example in medicine, isotopes with a short half-life are required (the half-life of technetium-99m, used as a radioactive tracer in the body, is about 6 hours). In contrast, isotopes used in radioactive dating need to have a relatively long half-life (the half-life of carbon-14 is 5700 years, but sometimes even that is not long enough; the half-life of uranium-238 is 4.5×10^9 years).

An analogy may help you to understand the random nature of radioactive decay and the concept of half-life. If you take a coin and toss it, the probability of getting 'heads' is 1/2, i.e. each time you toss the coin you have a 50 : 50 chance of this outcome. You have absolutely no way of knowing whether you'll get 'heads' or 'tails' on the next throw (in fact it would be possible to toss the coin 100 times and get

'tails' every time), but if you do a very large number of throws you'll get heads about half of the time. Now imagine that instead of tossing a coin you throw a die (note that 'die' is the singular of 'dice'). There are more options available than there were for the coin (six possible scores as opposed to just two — heads and tails), so each possible outcome (e.g. throwing a six) is less likely. So, although you never know whether a particular toss of a coin or throw of a die is going to produce heads or a six, it seems reasonable to suppose that a model of radioactive decay based on scoring a six when you throw a die will somehow be a slower process than a similar model based on getting heads when you throw a coin. You will investigate this further in Activity 3.3.

Activity 3.3 Tossing coins and throwing dice

We provided a group of Open University students with a tub containing 100 coins and asked them to do the following

- 1 Throw all the coins out of the tub onto a table.
- 2 Collect all the heads and count them, and *put these coins to one side*. These coins represent the number of atoms that decayed in the first time interval.
- 3 Return the other ('non-decayed') coins to the tub and throw again.
- 4 Again collect heads, count them and put these coins to one side.
- 5 Repeat until all the coins have been put to one side.

46 heads were collected on the first throw, 20 on the second, 17 on the third and so on. This simulation was labelled 'Run A' and the students then repeated the whole simulation a further nine times (runs B to J), and added the results together, so it was as if they had started with 1000 coins. Figure 3.12 shows the experiment in progress and the results are given in Table 3.3.



(a)



(b)



(c)

Figure 3.12 A simulation of radioactive decay using coins.
(a) The coins are thrown out of the tub (b) Heads ('decayed atoms') are counted and put to one side
(c) The other ('non-decayed') coins are returned to the tub.

Table 3.3 A simulation of radioactive decay using coins.

Throw number	Number of heads thrown										Total for runs A–J
	Run A	Run B	Run C	Run D	Run E	Run F	Run G	Run H	Run I	Run J	
1	46	50	55	51	43	53	54	53	58	48	511
2	20	24	19	20	26	33	23	27	25	32	249
3	17	14	9	15	21	2	8	11	5	8	110
4	8	6	9	7	8	6	7	6	6	8	71
5	5	3	5	1	0	4	6	3	2	2	31
6	3	2	3	4	1	1	2	-	0	2	18
7	0	1	-	1	1	0	-	-	4	-	7
8	0	-	-	0	-	1	-	-	-	-	1
9	0	-	-	1	-	-	-	-	-	-	1
10	0	-	-	-	-	-	-	-	-	-	0
11	1	-	-	-	-	-	-	-	-	-	1

Task 1

Plot points on the graph paper given in Figure 3.13 to show the total number of heads collected at each throw against the throw number. We have labelled the axes and plotted the first two points to help you.

Task 2

Draw the 'best fit' curve. Your points will not lie exactly on a smooth curve, but your best-fit curve should be a smooth curve that is as close as possible to as many of the points as possible.

Task 3

- Use your graph to find the number of throws required for the 'number of heads thrown' to halve. This represents the half-life of the process. Note that half-life does not depend on the actual number of coins (or atoms) decaying. You should get a similar result by considering, for example, the interval between 400 heads thrown and 200 heads thrown as you would get by considering the interval between 200 heads thrown and 100 heads thrown – the important thing is that the number of heads thrown is halved. Find three values for the half-life of the process from your graph, and then find their mean.
- Is the result what you expected or not?
- How would you expect a graph drawn for a very large set of data (say 10 000 coins to start with) to compare with the one you have drawn?
- Figure 3.14 gives results from a different simulation of radioactive decay, this time an idealized representation of the throwing of 1000 dice. The dice are considered to have 'decayed' when a six is thrown. How do the graph shown in Figure 3.14, and the value of the half-life for the process it represents, compare with the graph that you plotted for coins in Tasks 1 and 2 and the half-life that you found in Task 3(a)? ◀

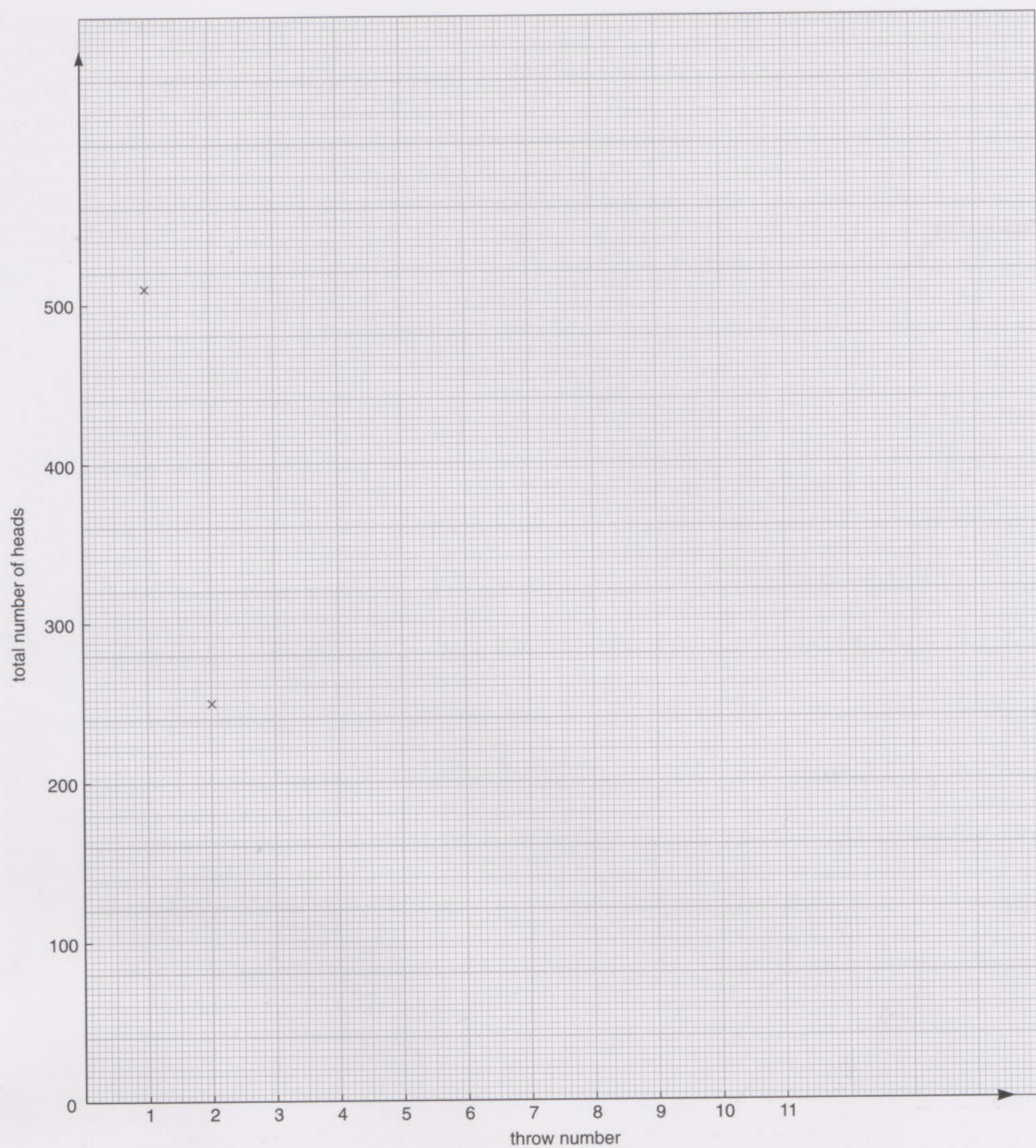


Figure 3.13 Graph paper for use in Activity 3.3.

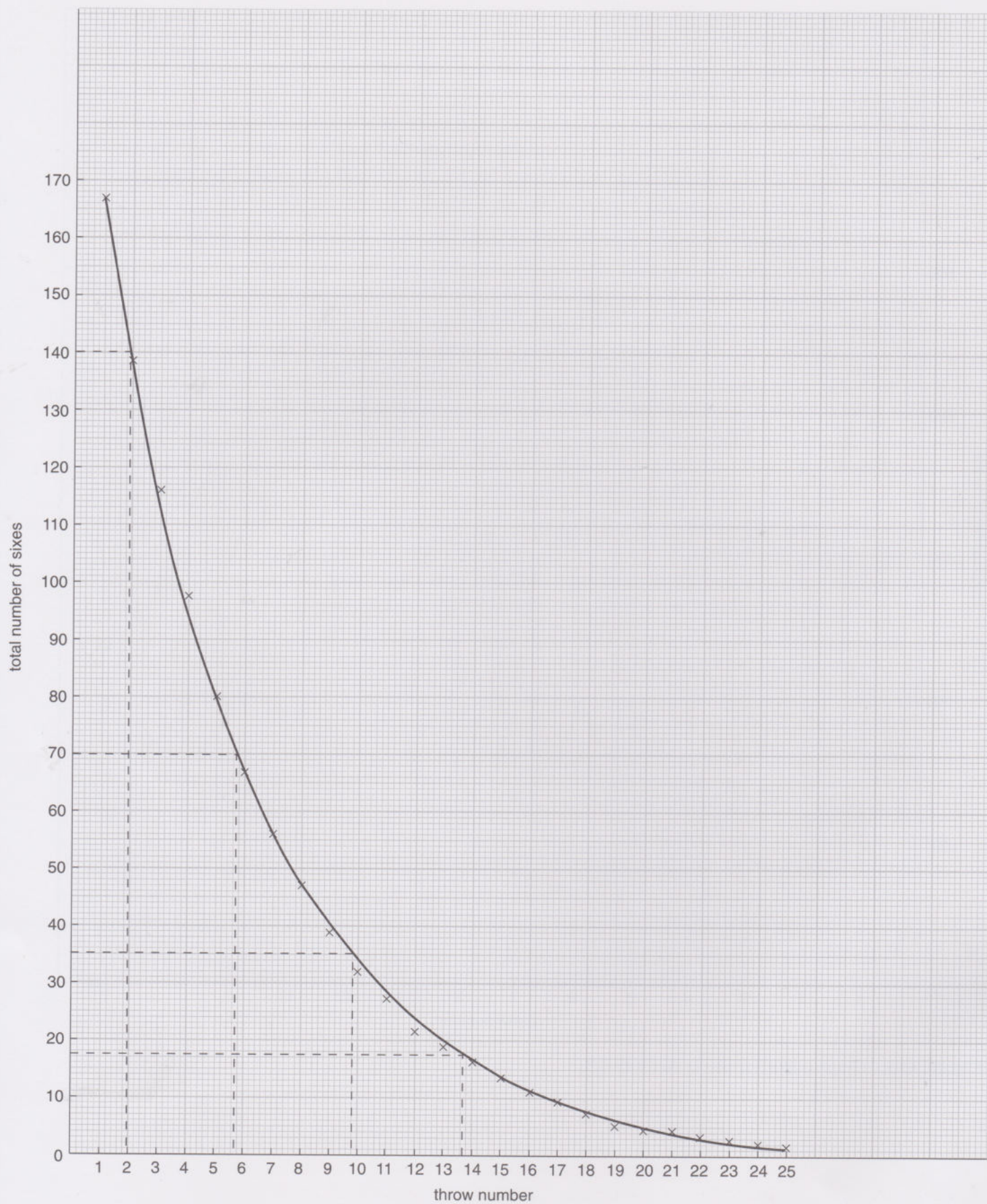


Figure 3.14 The total number of sixes thrown at each throw number.

The eagle-eyed amongst you may have spotted that in the coin and dice throwing simulations you are counting the number of heads and sixes thrown, with the analogy that a head or six corresponds to the decay of an atom, whereas the half-life for a radioactive decay process is described in terms of the number of undecayed atoms remaining, not the number of atoms decaying at any time. Fortunately there is no need to worry — the real radiation that you ‘count’ in the experiment at Residential School is a measure of atoms decaying, but the number of atoms decaying at any time is proportional to the number of undecayed atoms remaining, so the half-life is identical however it is calculated.

3.5.6 ‘Counting’ radioactivity

In Activity A you will be using a number of different radioactive sources, and you will ‘count’ the radioactivity that they produce by means of a Geiger counter^G of the type shown in Figure 3.15. The cap has been removed from the end of the Geiger tube in Figure 3.15 so that you can see the entrance window through which radiation passes. When an α -particle, an electron or a γ -ray photon enters the Geiger tube, it creates an electrical signal that is recorded by the attached counter. The digital display on the counter records the number of radioactive ‘particles’ that have been detected. The counter can also be set up to ‘click’ every time a particle has been detected.



Figure 3.15 A Geiger counter, consisting of a tube connected to a meter that displays the measurements.

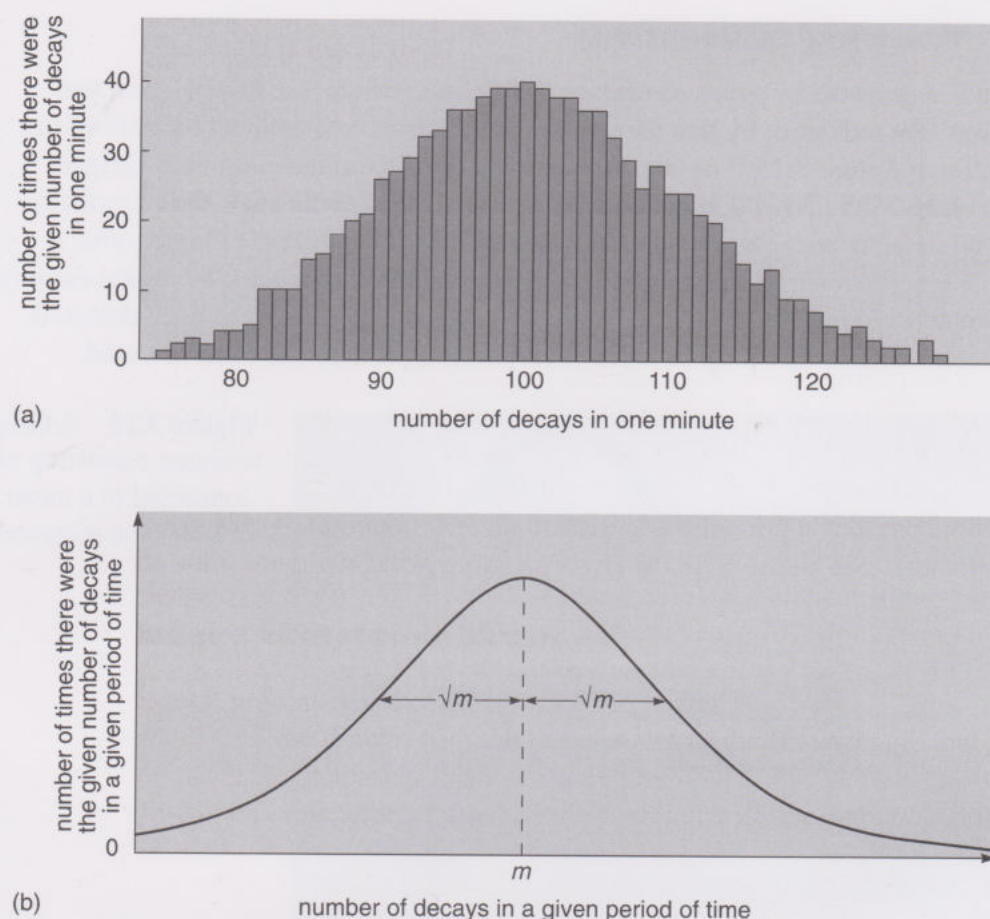
3.5.7 Counting statistics and uncertainties

If you look back to Table 3.3 you will see that there was considerable variation between the results of the different runs in the tossing coins simulation, for example in Run C, 55 heads were collected for the first throw whereas in Run E only 43 heads were collected for the first throw. This is because the process is governed by probability, so we can never know how many heads are going to be thrown. However, as more and more coins are tossed, the overall ratio of heads to tails becomes closer and closer to the theoretical 50 : 50. A similar situation applies every time you use a Geiger counter to ‘count’ radioactive decays. If you use the same Geiger counter at a fixed distance from a fixed source to take the five consecutive readings for the count in the same period of time, your readings are likely to be different. However, if you were able to repeat the measurement, say, 1000 times, a pattern would begin to emerge. Figure 3.16a shows a histogram of 1000 such measurements, plotted as the

'number of times there were the given number of decays in one minute' against the 'number of decays in one minute'. As you can see there is a most probable value for the number of decays, which in this case is about 100 decays in one minute.

The overall shape displayed by the histogram in Figure 3.16a can be approximated by a well-known mathematical curve which is drawn in Figure 3.16b. Don't worry too much about the detail of Figure 3.16, or the following two paragraphs of text. The important thing is that you know how to calculate the uncertainty in a count rate; not that you understand the underlying theory.

Figure 3.16 (a) A histogram showing the results of repeating a background count measurement one thousand times. (b) The mathematical curve to which the histogram may be approximated.



This curve shown in Figure 3.16b has two useful properties. First, the position of the peak of the distribution corresponds to the mean value m of all the measurements: the mean value is therefore also the most probable value. Second, 68% of all the measurements lie within $\sqrt{m}$ either side of this mean ($\sqrt{m}$ means the square root of m and square roots are discussed on page 356 of *SGSG*). So there is a 68% probability that any individual measurement will lie within $\sqrt{m}$ of the most probable value of m , i.e. $m \pm \sqrt{m}$. For the results shown in Figure 3.16a, since the square root of 100 is 10, this means that 68% of the measurements lie between $(100 - 10) = 90$ and $(100 + 10) = 110$ decays in one minute, or 100 ± 10 .

Now, suppose you haven't the time (or patience!) to take 1000 measurements, but you take *just one* measurement with a value n , say. Then there is a 68% probability that the single measurement you have taken lies within $\sqrt{m}$ either side of the mean value m that you would get if you took 1000 measurements. You don't know what this mean value is (you only have the one measurement), but we can turn the argument round and say that there is a 68% probability that the mean value of a thousand measurements would lie within $\sqrt{n}$ of the single measurement with value n .

This will save you a lot of time! Rather than taking 1000 measurements, just take one measurement with a value n and then assume that there is a 68% probability that the mean value of 1000 measurements would lie somewhere within the range covered by $n \pm \sqrt{n}$. 68% may not seem like a very high probability, but it is perfectly acceptable in practice. The 'plus or minus the square root of n ' is referred to as the uncertainty^G of the particular value n that you have measured. In general, any radioactive count that you record with a Geiger counter should be expressed in the form of the number n with its associated uncertainty, given by $\pm \sqrt{n}$.

● Suppose you use a Geiger counter to obtain a count rate of 57 counts per minute. Give this count rate with its associated uncertainty (round the uncertainty up to the whole number above it).

● $\sqrt{57} = 7.55 = 8$ to the whole number above, so the count rate per minute is 57 ± 8 .

You will need to think about experimental uncertainties frequently when you are at Residential School. In general, the uncertainty in a measurement is an estimate of its true range. We have already discussed the need to give uncertainties with your readings for count rate in Activity A. In Activity B, 'Analysing our environment', you will discover that it is only possible to measure the angles of diffraction on your spectrometer to the nearest 0.1° , so the uncertainty is $\pm 0.1^\circ$. This leads to an uncertainty of ± 0.001 or ± 0.002 in the sines of the angles (you will learn how to calculate these values at Residential School) and you can then calculate the uncertainty in your calculated values for grating spacing, d and wavelength, λ . In Activity D, 'The biological effects of gamma radiation', the measurements that you obtain for the lengths of wheat seedlings will have, in addition to measurement uncertainties (because you can only measure to a certain precision), a range of values caused by the inherent variability of the seedlings.

Uncertainties in values can be shown on graphs by means of error bars^G. The value of count rate of 57 ± 8 counts per minute given above would be plotted as shown in Figure 3.17. The central point is at 57 counts per minute. The upper end of the error bar is plotted at $57 + 8 = 65$ counts per minute and the lower end is plotted at $57 - 8 = 49$ counts per minute. The line drawn between the two points spans the range of values consistent with the measured quantity.

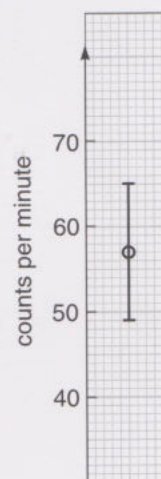


Figure 3.17 A count rate of 57 ± 8 counts per minute plotted on a graph, with the uncertainty shown as an error bar.

3.5.8 Background radiation

A Geiger counter will always detect some radiation whether or not it is pointed at a sample of radioactive material. This is because so-called background radiation^G is around us all the time.

- About 500 000 cosmic rays from outer space penetrate the average person every hour.
- About 30 000 atoms of radioactive radon, polonium, bismuth or lead in the air we breath disintegrate each hour in an average person's lungs.
- About 15 million atoms of potassium-40 and 7000 atoms of uranium, from the food we eat, disintegrate per hour inside an average person.
- Over 200 million γ -rays from the soil and buildings pass through an average person each hour.

3.5.9 Radiation safety

The radioactive sources that you will use in Activity A are very weak. The strongest source produces about the same number of radioactive decays as the radioactive source in a household smoke detector, namely 40 kilobecquerels^G (40 000 radioactive decays per second).

We have determined that the maximum dose equivalent^G of radiation that you might receive during Activity A, as a result of being near the radioactive sources, is less than 1 microsievert (μSv). The definition of the unit of dose equivalent, the sievert^G (Sv), is complex, depending on the absorbed dose in the time you are close to the source and the different biological properties of different types of radiation and particles, but the following comparisons may be of interest:

- The typical dose equivalent experienced as a result of natural background radiation in the UK is about $6.8 \mu\text{Sv}$ per day, or $21 \mu\text{Sv}$ per day if you live in a location such as Cornwall with a high concentration of radioactive rocks.
- On a 3-hour aeroplane flight you would receive a dose equivalent of about $10 \mu\text{Sv}$ from cosmic rays.
- The dose equivalent that you would receive from a typical chest X-ray is about $20 \mu\text{Sv}$.
- It is estimated that the instantaneous dose of γ -radiation to which some of the survivors of the Hiroshima and Nagasaki atomic bombs were exposed was approximately 5 Sv.
- As stated above, we estimate that you will receive a dose equivalent of about $2 \mu\text{Sv}$ in performing the experiments in Activity A.

Question 3.3 Calculate how many times greater the dose equivalent mentioned above as being received by survivors of the Hiroshima and Nagasaki atomic bombs is than the dose equivalent that you might receive during Activity A.

In Activity D you will be studying the effects that γ -radiation has on the growth of a species of wheat. Wheat grains for this experiment are exposed to 2 kSv of γ -radiation before delivery to the Residential School site.

$$\frac{2 \text{ kSv}}{2 \mu\text{Sv}} = \frac{2 \times 10^3 \text{ Sv}}{2 \times 10^{-6} \text{ Sv}} = 10^9$$

so this is 10^9 (a thousand million) times greater than the dose equivalent that you might receive in Activity A. Note, however, that neither the wheat grains nor the plants subsequently grown from them are made radioactive by this treatment and both can be handled with complete safety.

3.5.10 The effect of distance

We are now in a position to be able to explain some of the safety precautions detailed in the Activity A workbook (which you might like to look at now), for example, the instructions to use tongs when handling radioactive samples and to keep samples at the back of your bench when not in use. If you were to use a Geiger counter to count the radiation 1 cm from a point source and then repeat the readings at a distance of 10 cm from the same source, you would find that the count rate would be reduced by a factor of approximately 100. The variation of count rate with distance obeys an inverse square law^G, in other words the count rate decreases by a factor of x^2 every

time the distance increases by a factor of x . So if you use tongs to handle the uranium-rich mineral encased in resin, which is illustrated in Figure 10 of the workbook for Activity A your hand will be at a distance of approximately 20 cm from the mineral rather than at a distance of about 1 cm, so the dose of radiation which you receive will be reduced by a factor of $(20)^2$ i.e. 400.

The marked variation of count rate with distance also explains why it is so important to keep the positions of the Geiger tube and the isotope generator *fixed* when you are investigating the absorption of γ -rays by different thicknesses of lead.

3.5.11 The absorption of radioactivity

In Activity A, you will be determining the half-life for a radioactive source, but you will also be investigating what happens when α - or β -particles or γ -rays pass through different materials. We can quantify the effectiveness of a material in absorbing γ -rays by stating its half-thickness^G. If we plot a graph of count rate against the thickness of lead, say, between a fixed source and a fixed Geiger counter, we obtain a graph with the same characteristic shape as is obtained in calculating half-life, and the half-thickness is defined as the thickness of lead needed for the count to drop by half.

For γ -rays, the half-thickness depends only on the material used and the energy of the source. As the γ -rays pass through the lead they are scattered by the electrons in the metal, transferring some of their energy to the electrons with each encounter. The more electrons there are, the more likely it is that the energy will be transferred, so dense materials with a high atomic number (such as lead with an atomic number of 82) are relatively good absorbers of γ -rays. The net result therefore is that the energy of the γ -ray photons is transferred to the lead.

The absorption characteristics of α -particles and electrons are different from those of γ -rays. Alpha-particles are the most readily absorbed, being completely stopped by a single sheet of paper, whereas electrons are absorbed easily by low atomic number materials, such as the carbon atoms in Perspex. Neither of these materials is very effective at stopping γ -rays. Lead is an effective absorber of all types of radiation, which is why it is used to shield people from radiation sources in 'lead aprons' etc., and also in the boxes used to store the radioactive sources that you will use at Residential School.

Gamma-rays, α -particles and electrons released in the Earth from radioactive minerals are absorbed by the minerals themselves and the surrounding rocks. Figure 3.18 illustrates the radiation damage that can result. The energy released by radioactive decays in the Earth can be transferred to the Earth itself. You will investigate the amount of energy that is transferred in this way in the final part of Activity A.

Figure 3.18 Radiation damage in and around a mineral containing uranium.



3.6 Self-assessment questions

Question 3.4 The mass of a proton (m_p) is 1.673×10^{-27} kg and the mass of an electron (m_e) is 9.11×10^{-31} kg. Write the mass of the electron as a fraction of the mass of the proton.

(Express your answer in the form $m_e = (1/n)m_p$, where n is a whole number quoted to an accuracy of 3 significant figures.) ◀

Question 3.5

(a) A vapour of substance X is heated and gives out (emits) visible light that comprises a spectrum of a number of coloured lines (each line corresponding to a specific wavelength). Explain, in a few sentences, the link between this spectrum and the processes taking place in the atoms of substance X.

(b) In the hydrogen spectrum shown in Figure 3.6b, the red hydrogen spectral line is caused by electrons falling from energy level 3 to energy level 2. If the energy difference $(E_3 - E_2) = 3.03 \times 10^{-19}$ J, what is the wavelength of the red line?

(Assume values of 6.63×10^{-34} J s for the Planck constant, h , and 3.00×10^8 m s⁻¹ for the speed of light in a vacuum, c) ◀

Question 3.6 Light from a laser, with a wavelength of 633 nm, is shone (at 90°) on to a diffraction grating that has an unknown grating spacing. The angle between the straight through direction and the first order diffraction image is 30.4°. Use this information to calculate the value of the grating spacing. ◀

Question 3.7 A student measures the speed of a falling stone at time intervals of 0.5 second. The data collected are recorded in Table 3.4. (The uncertainty in the time is negligible compared with the uncertainty in the speed measurement.)

Using these data, draw a graph of speed against time on the graph paper provided (Figure 3.19), remembering to include error bars to show uncertainties.

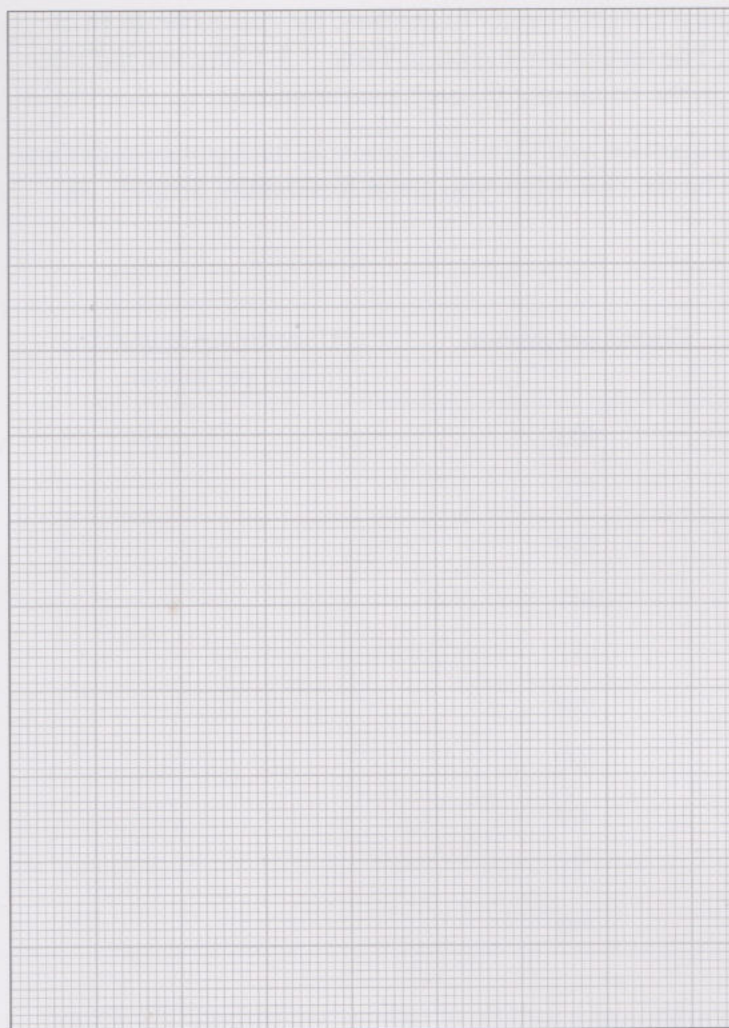


Figure 3.19 Blank graph paper for use with Question 3.7.

The gradient of the graph gives the average acceleration of the stone. Use your graph to find this acceleration. ◀

Table 3.4

Time/s	Speed of stone/m s ⁻¹
0	
0.5	5 ± 1
1.0	10 ± 1
1.5	15 ± 1
2.0	20 ± 1
2.5	26 ± 2
3.0	29 ± 2

Question 3.8 A student used a Geiger counter to measure the count rate from an unknown radioactive source. The results are shown in Table 3.5. Complete the table by including the uncertainty in the count rate for each measurement.

N.B. For the purposes of this question you should assume that background radiation can be ignored. This will *not* be the case when you are at Residential School. ◀

Table 3.5

Elapsed time/min	Counts per minute	Uncertainty in count rate per minute
1	403	
2	320	
3	249	
4	201	
5	157	
6	128	
7	109	
8	89	
9	71	
10	66	

Question 3.9 Figure 3.20 shows how a sample of radioactive carbon $^{14}_6\text{C}$ decays from its natural state. Use this graph to find the half-life of $^{14}_6\text{C}$. ◀

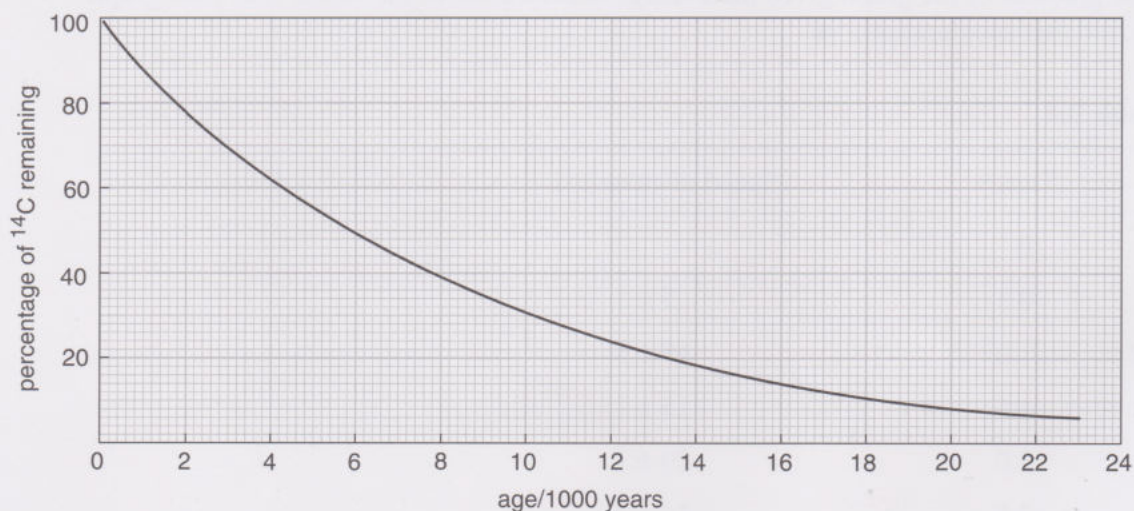


Figure 3.20 The radioactive decay curve of $^{14}_6\text{C}$.

3.7 What to do next

You should now read the second half of the Activity A workbook (from Task 8 onwards). You should also read Section 3 of Chapter 8 of *SGSG* starting on page 204.

3.8 Summary of Section 3

An atom is a very small structure that is about 10^{-10} m across. It consists of a nucleus, containing protons and neutrons, surrounded by a distribution of negatively charged electrons.

Energy is conserved: thus when an atom changes from a state of higher energy to one of lower energy, a photon with energy corresponding to the atom's loss of energy is emitted.

Different elements have different spectral signatures, consisting of spectral lines of characteristic energies and wavelengths.

A spectrometer can be used with a diffraction grating to produce a diffraction pattern. The angle of diffraction θ_n for the n th order is related to the grating spacing d and the wavelength λ by the equation

$$\sin \theta_n = \frac{n\lambda}{d}$$

Some chemical elements exist in several forms, known as isotopes, each with the same atomic number (i.e. the same number of protons) but with a different number of neutrons and hence a different mass number.

Radioactive decay is the process whereby certain unstable nuclei break up or decay spontaneously. The types of decay include α -decay, β^- -decay and γ -decay. In any nuclear decay, electric charge and mass number are always conserved.

An infinitesimal amount of mass, m , is lost in radioactive decays, and this is converted into energy, E , according to the equation $E = mc^2$, where c is the speed of light in a vacuum.

Radioactive decay is an intrinsically random process and we can never know when an individual nucleus will decay. However, if we take a large enough number of nuclei, we can use the concept of half-life (the time taken for half a radioactive sample to decay) as a measure of how rapidly the isotope is likely to decay.

A Geiger counter will always detect some radiation whether or not it is pointed at a source of radioactive material. This is because of background radiation, which is around us all the time.

Alpha-particles, beta-particles and gamma rays are absorbed as they pass through materials. We can measure the effectiveness of a particular type of particle or radiation by quoting its half-thickness.

There is a degree of uncertainty in all experimental measurements. Uncertainties can be indicated on graphs as error bars.

The radioactive sources you will use at Residential School are very weak. Additional safety precautions are based on keeping them at a distance from your body (e.g. by using tongs) and by storing them in lead-lined boxes when not in use, as lead is a very effective absorber of all types of radiation.

Now that you have completed Section 3 you should be able to:

- explain the link between the lines observed in emission spectra and the processes that take place within atoms;

- calculate the grating spacing, d , of a diffraction grating from knowledge of diffraction order, n , angle of diffraction, θ_n , and wavelength, λ ; or calculate an unknown wavelength from knowledge of d , n and θ_n ;
- plot a best-fit graph and calculate and interpret the gradient of a straight-line graph;
- calculate the uncertainty in a measured value of count rate;
- calculate half-life or half-thickness from an appropriate graph of count rate against time or thickness of absorber;
- explain why it is important to keep radioactive sources at a distance from your body, and to store them in a lead-lined container;

If you have experienced difficulties with the mathematics in this section, you may find the Section 'Maths Help' in *SGSG* useful (page 302 of *SGSG* lists the areas covered).

Biology: the science of life

4

Biology features prominently in two of the five activities that comprise the residential component of *Practising Science*: Activity C ‘Investigating the environment’ and Activity D ‘The biological effects of gamma radiation’. Activity C integrates Earth sciences and ecology (the branch of biology concerned with the interaction of living organisms with their environment, including other living organisms) in the context of fieldwork. In contrast, Activity D is a laboratory-based investigation predominantly biological in nature. Although they are concerned with rather different areas of biological knowledge and employ quite different practical techniques, Activities C and D do have several important features in common. For instance, both require you to make careful observations (both qualitative and quantitative) of living organisms, to record data systematically, and then to propose and test hypotheses that might explain your initial observations.

To ensure that you have the necessary biological knowledge and skills to get the most out of these activities, you should first work through Section 4.1 (which supports Activity D) and then Section 4.2 (which supports Activity C).

4.1 Activity D ‘The biological effects of gamma radiation’

The work that you are expected to do in connection with Activity D *prior* to the residential component of *Practising Science* is of two sorts. First, you must ensure that you are familiar with (i) the basic structure of plant cells (Section 4.1.1) and (ii) the process of cell division (including the details of mitosis) (Section 4.1.2). The terms whose meaning Residential School tutors will assume that you understand — and which are used without explanation in the Activity D workbook — are given the superscript ‘G’ here (to indicate that they are defined in the Glossary). Second, you should carry out Part I of Activity D itself (which is located in the workbook) (Section 4.1.3). Part II (the work that you will be doing in the laboratory) builds directly on Part I.

4.1.1 Structure of plant cells

Since plants are multicellular^G eukaryotes^G (as opposed to prokaryotes, such as bacteria), their cells^G consist of cytoplasm^G (surrounded by a cell membrane^G) and a cell nucleus^G (surrounded by a nuclear membrane^G). Two additional features possessed by plant, but not animal, cells are external cell walls^G made of cellulose (which give the cells their shape and rigidity) and, within the cytoplasm, chloroplasts^G (responsible for photosynthesis^G) (Figure 4.1).

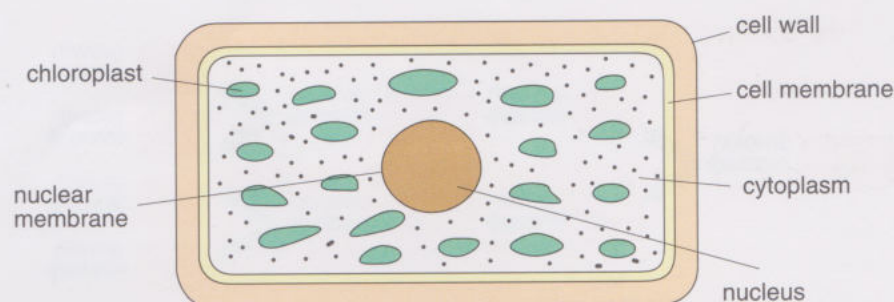


Figure 4.1 Schematic diagram of a typical plant cell. Actively photosynthesizing cells contain *many* more chloroplasts than shown here.

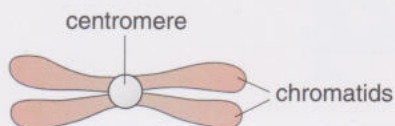


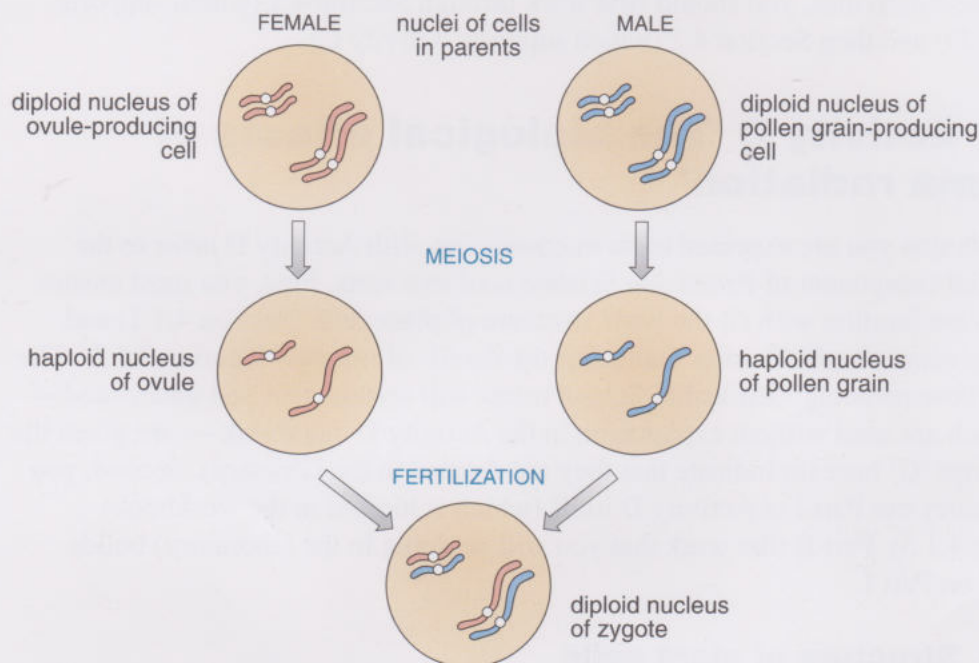
Figure 4.2 Sketch of a single chromosome to show the two chromatids joined at the centromere.

4.1.2 Cell division

The plant's genetic information is encoded in DNA^G, most of which is located in the nucleus. Nuclear DNA is packaged into chromosomes^G, which — most of the time — comprise two chromatids^G joined at a point known as the centromere^G (Figure 4.2). Each chromatid consists of a single relatively long (but extremely thin) DNA molecule with associated protein molecules. Between cell divisions — in what is called interphase^G — individual chromosomes cannot be distinguished within the nucleus even under a microscope.

A plant produced by sexual reproduction starts as a zygote^G formed by the fusion of two gametes^G at fertilization (i.e. an ovule from the female parent and a pollen grain^G from the male parent). Gametes have half the number of chromosomes of the zygote (i.e. they are haploid^G). The zygote and all the cells descended from it are described as diploid^G. Gametes are produced by meiosis^G, a process that — among other things — halves the number of chromosomes in the resulting progeny cells^G. Figure 4.3 shows how the number of chromosomes changes during meiosis and fertilization.

Figure 4.3 Diagram to show how the number of chromosomes changes during meiosis and fertilization for a hypothetical plant with only two pairs of chromosomes (most species have many more). For simplicity, each chromosome is shown as a single strand, rather than as a pair of chromatids.



All the plant's cells other than its gametes are produced by mitosis^G, a process that results in two cells that are identical with the parent cell — and hence one another — in all respects (except for relatively infrequent mutations^G).

The zygote grows in size, undergoes mitosis and then divides into two progeny cells. This sequence (known as the cell cycle^G) is repeated many times (Figure 4.4) to produce first an embryo, then a seedling and ultimately a complete flowering plant.

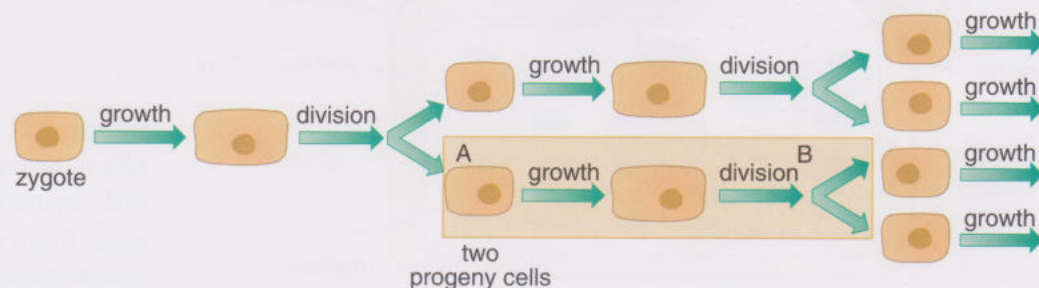


Figure 4.4 Repetition of the cell cycle (i.e. alternation of cell growth and mitotic cell division) enables the zygote to grow into a complete flowering plant.

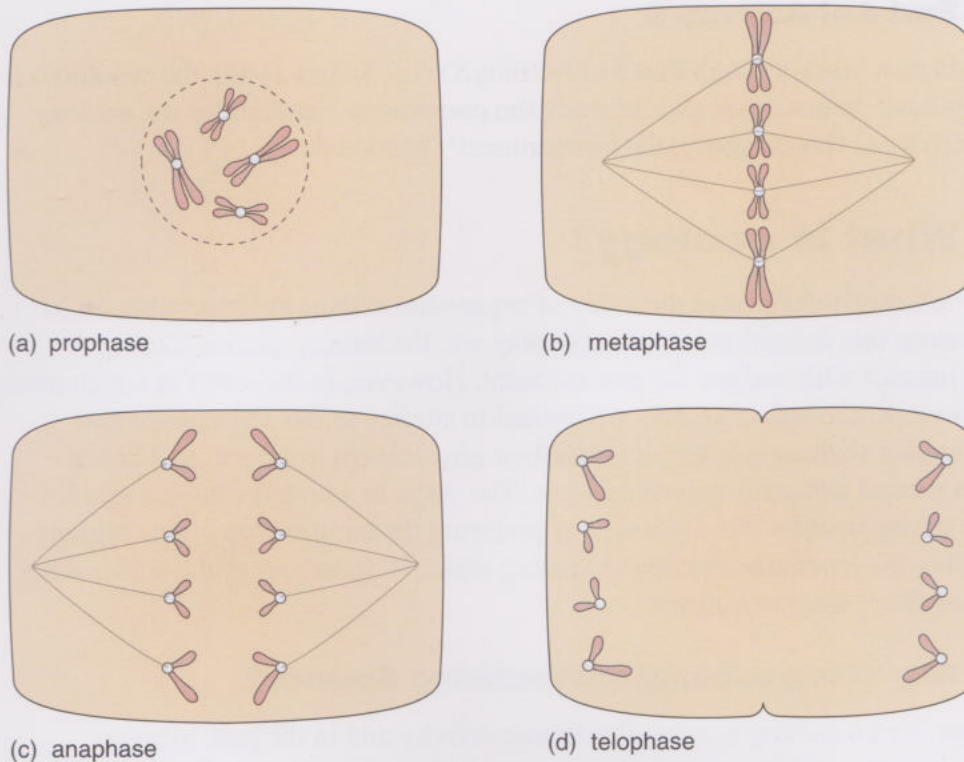


Figure 4.5 The four main phases of mitosis: (a) prophase; (b) metaphase; (c) anaphase; and (d) telophase (with cell division itself just starting).

Mitosis is a continuous process that takes several hours to several tens of hours (depending on the species and the prevailing conditions), but conventionally it is divided into four main phases (Figure 4.5).

During prophase^G the nuclear membrane disappears and, by coiling of their DNA molecules, the chromosomes become so short and fat (compared with their condition during interphase) that they can readily be seen under a microscope (provided suitable stains are used to make them visible) (Figure 4.5a). Each chromosome's centromere then connects to delicate threads attached at each end of the cell. These threads exert tension on the chromosomes so that they become aligned across the middle (or 'equator') of the cell (metaphase^G) (Figure 4.5b). The chromatids then separate (by splitting of the centromere), so that each becomes a chromosome in its own right. One member of each former pair of chromatids is drawn to one end of the cell, while its partner is drawn to the other end (anaphase^G) (Figure 4.5c). Once the chromatids reach one or other end of the cell, the threads that were attached to them disappear. There is now a set of single-chromatid chromosomes at one end of the cell and an equivalent set at the other end (telophase^G) (Figure 4.5d).

The DNA molecules then uncoil again, causing the chromosomes to disappear from view. At the same time, a nuclear membrane forms around each cluster of chromosomes so that the cell temporarily contains two nuclei. Mitosis is now complete, but the cell itself must still divide into two (i.e. each nucleus and its surrounding cytoplasm must become completely surrounded by a cell membrane and a cell wall). The two progeny cells — each of which contains an identical copy of the parent cell's genetic material — are now in interphase (as in Figure 4.1). Before the progeny cells can undergo mitosis, their DNA molecules must be replicated (i.e. their single-chromatid chromosomes must become two-chromatid chromosomes; Figure 4.2).

4.1.3 Part I of Activity D

You should now work through Part I of Activity D (i.e. Stages 1–4 in the workbook). When you have done so, you should study the preparatory material for the ecology part of Activity C ‘Investigating the environment’ (Section 4.2).

4.2 What is ecology?

Ecology^G is usually defined as the study of organisms in their environments. In its broadest sense this definition includes the way we, the human species (*Homo sapiens*), interact with and use the environment. However, in the sense in which most ecologists work, ecological studies are limited to studies of the ways plants and animals interact with each other and with their physical environment, and hence survive in natural and semi-natural habitats. The ways in which the human species affect the environment — for instance, by polluting the air, growing crops, digging peat, eroding the landscape or even extracting water — form part of the study of the factors that affect other organisms.

4.2.1 Why study ecology in *Practising Science*?

These days, bird watching is a popular leisure activity and in the past so were collecting insects, wild flowers and birds’ eggs (although such activities are not now recommended — indeed, they are often illegal — because of the potential damage they cause to flora^G and fauna^G). Some amateurs are or were truly experts in their fields. In fact, much of the original identification of the British flora and fauna was done by amateur naturalists. Many a Victorian vicar or other self-taught naturalist contributed to the specialist knowledge available to, and used by, professional biologists today. And though ‘high-tech’ techniques are often used by professionals, many aspects of ecology can be studied with comparatively simple and limited equipment — provided one has time and patience.

Stories of environmental damage appear fairly frequently in the national press. Peat extraction (gardeners are now encouraged to use composts that do not have a peat base) and pollution of air and water are frequently highlighted. Such events have marked effects on local plant and animal populations by killing off organisms, or in the case of peat extraction, by eliminating whole habitats from a local area. Is this important?

There are obviously aesthetic reasons for preserving threatened habitats; diversity of landscape forms is pleasing to the eye — for instance, that created by hedgerows along roadsides and between fields, or by small woodlands. The general public usually appreciates this argument even when an ecologist’s view, that maintaining the diversity of organisms and habitat is important for its own sake, is not accepted so readily. Preservation of different habitats in the landscape is closely associated with maintaining biodiversity^G and the healthy perpetuation of the flora and fauna, including rare species.

In many cases reduction of numbers leading to the elimination of organisms is an indicator that ecologically something has gone badly wrong. In time, this may even adversely affect the human population.

Figure 4.6 shows the number of species of lichens^G and bryophytes^G (i.e. mosses and liverworts) growing on trees and on walls at different distances westward (i.e. towards the prevailing wind) from the centre of Newcastle upon Tyne. Table 4.1

gives the annual atmospheric sulfur dioxide (formula SO_2) concentrations along a transect^G north-west from Newcastle upon Tyne city centre at about the time that these data were collected.

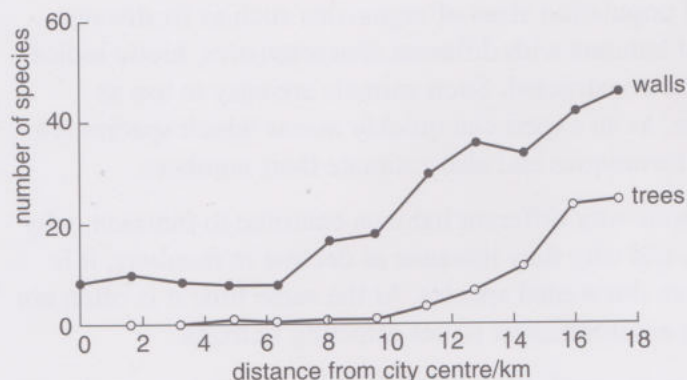


Figure 4.6 Numbers of species of lichens and bryophytes growing on trees and on walls at different distances westward from the centre of Newcastle upon Tyne.

Table 4.1 Annual atmospheric average SO_2 concentration along a transect north-west from the centre of Newcastle upon Tyne.

Distance from city centre/km	SO_2 concentration/ g m^{-3}
0	200
2.5	160
8	90
65	65

Question 4.1 From the information in Figure 4.6 and Table 4.1, describe the likely effect of SO_2 pollution of the air on the growth of lichens and bryophytes. ◀

Question 4.2 Similarly, extreme levels of water pollution are often accompanied by fish deaths (Figure 4.7). Do you think that other organisms were killed, or only fish? ◀

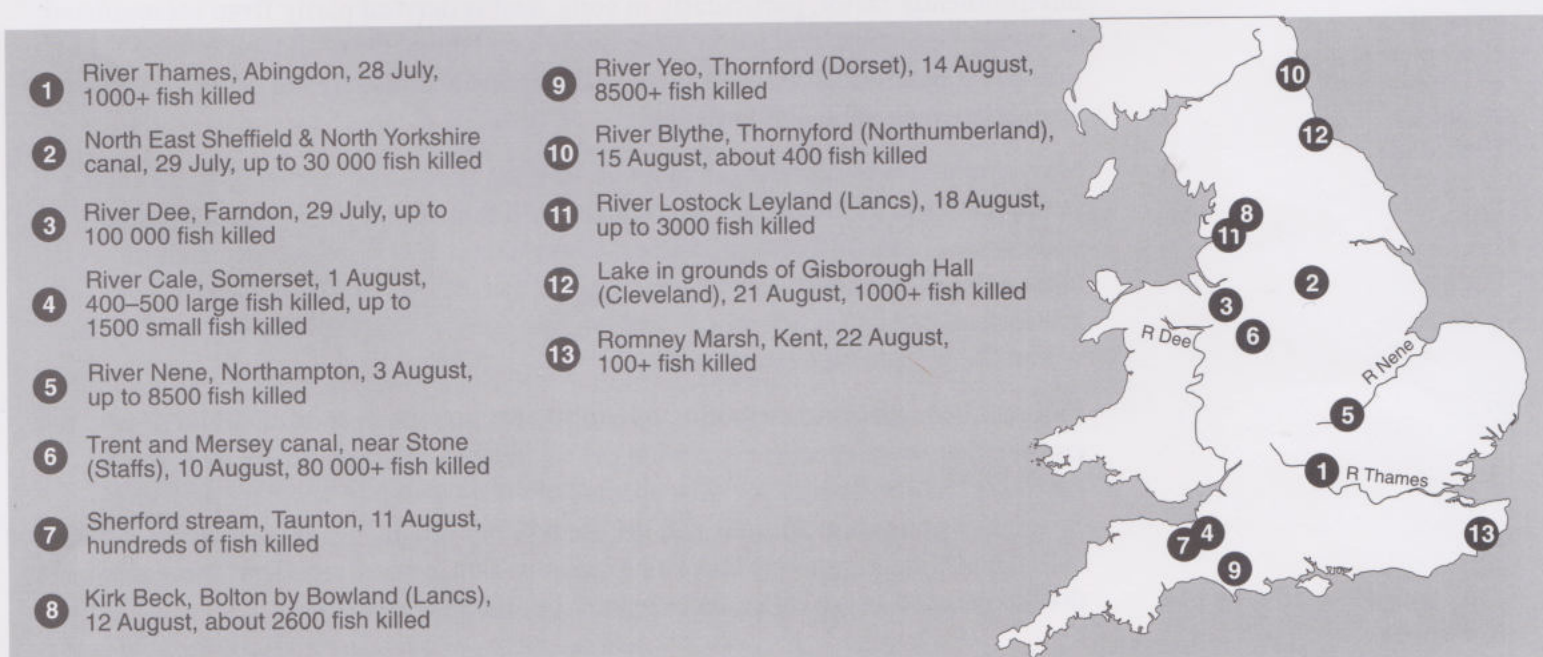


Figure 4.7 Water bodies in which there were severe water pollution incidents during the period from 28 July to 22 August 2000, and the numbers of fish killed.

With increasing levels of pollution — whether of air or of water — there is likely to be a gradual reduction in the population sizes, and eventually complete elimination, of species that are more susceptible to pollution.

Using the presence/absence and population sizes of organisms such as freshwater invertebrates^G that are typical of habitats with different characteristics, biotic indices measuring pollution levels can be constructed. Such animals are easy to use as indicators of the level of damage, as an expert can quickly assess which species are present using simple collection techniques and also estimate their numbers.

If we do not understand the reasons why different habitats continue to function, why organisms occur where they do, and why they increase or decline in numbers, it is impossible to identify or conserve threatened species. At the same time it is often not possible to make judgements on environmental issues affecting humans.

But why not leave such matters to the professionals? Why learn about how ecologists work? There certainly seems to be increasing concern about the soundness of much environmental decision-making. If you, as a member of the public, are to make valid judgements about information released in the press, you usually need some knowledge about the subject and especially about the ways in which data are collected and their reliability. This in itself is a good justification for including field ecology in the residential component of *Practising Science*.

4.2.2 Two factors affecting the distribution of organisms

The fieldwork you will do when you go to the Residential School will involve study of *either* two types of grassland (Sussex University) *or* a rocky sea-shore (Heriot-Watt University). We will illustrate some of the complexities of interpreting ecological field data by looking at two sets of factors, each of which is important in one of these habitats.

Soil pH

pH^G (a measure of acidity or alkalinity, discussed in Section 5.2) is an important environmental factor, particularly in soils. Soil is derived partly from accumulated decaying vegetation and partly from broken up fragments of the underlying rocks. Soil pH is determined by both these components and also by the water that fills the spaces between solid soil particles.

How might you expect the pH of soil overlying limestone^G (or chalk^G, which is a particular form of limestone) to compare with that of soil overlying sandstone^G? The soils would have different pH values. Limestone is a calcareous rock rich in carbonates^G (usually calcite^G), which gives rise to soils of pH of 7 and above. Sandstones contain much silica^G and so give rise to neutral or slightly acid soils, where the pH can be as low as 3.5.

Figure 4.8 shows the distribution of earthworm species in soils of different pH. You can see that some species have wide pH tolerances, whereas others have much narrower ranges, being specialist species of either acidic or more neutral soils. Species of plants and animals that are found only in alkaline environments are termed calcicoles^G (i.e. organisms that like a calcium-rich environment), whereas those that dislike these soils are called calcifuges^G (i.e. calcium escapers).

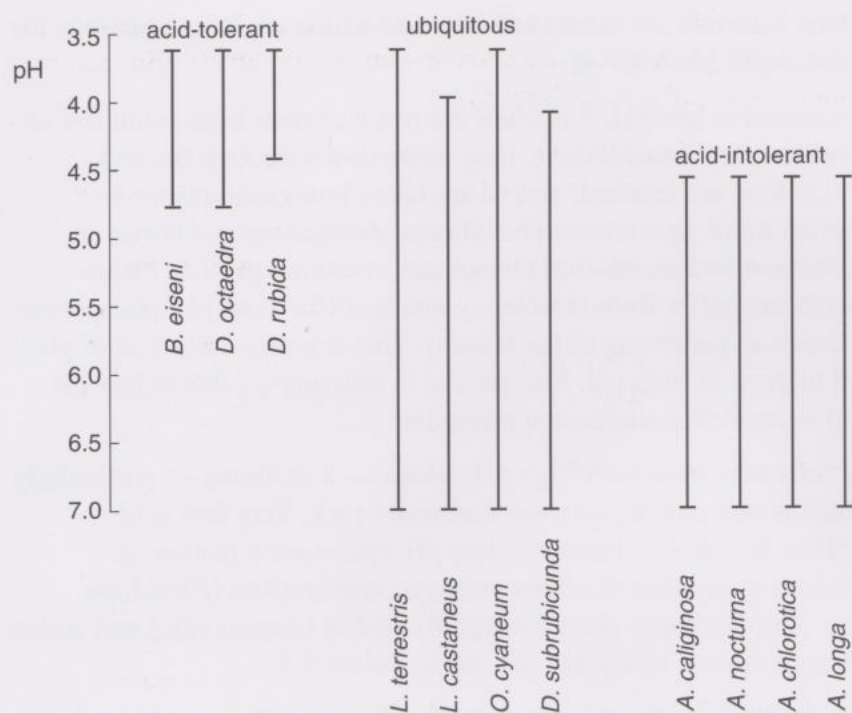


Figure 4.8 Distribution of earthworm species in soils of different pH. (*B.* = *Bimastos*; *D.* = *Dendrobaena*; *L.* = *Lumbricus*; *O.* = *Octolasmus*; *A.* = *Allobophora*.)

pH is a critical factor for plants, if only because it affects the solubility of mineral nutrients and hence their availability to plants. Figure 4.9 shows, in the form of 'kite' diagrams, the relative availability of several minerals at different pH values. A mineral is most available at pH values where its 'kite' is widest and least available where its 'kite' is narrowest or non-existent.

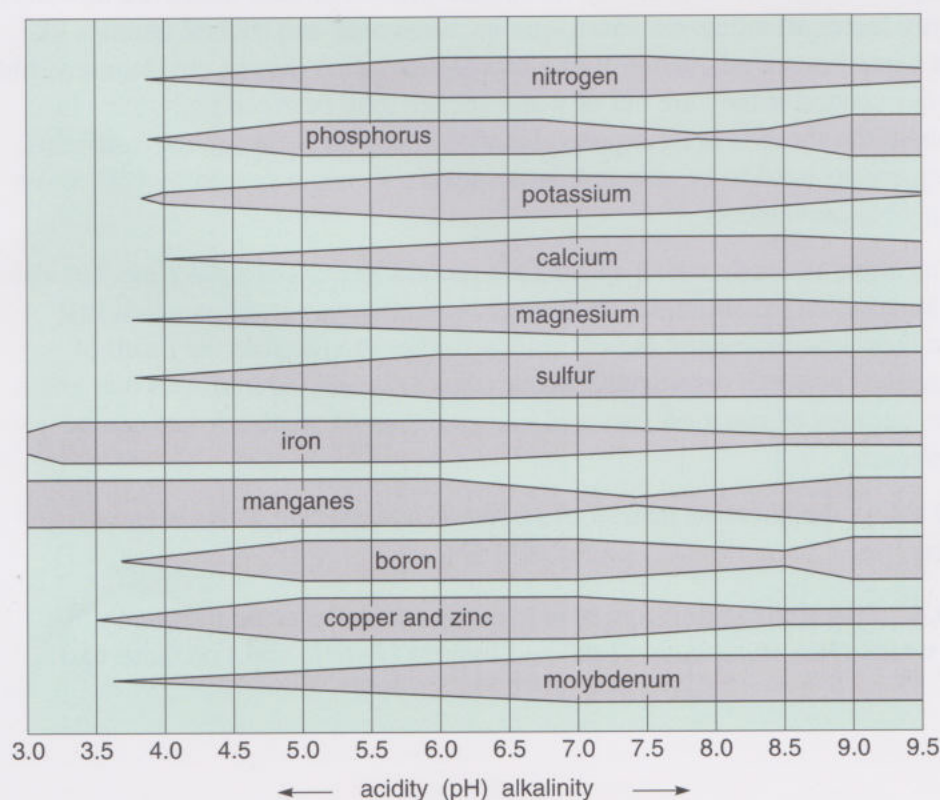


Figure 4.9 Availability to plants of some mineral nutrients in soils of different pH.

Question 4.3 Which minerals are most available, and which are least available, (a) at pH 7.5–8.5 and (b) at pH 3.5–4.5? ◀

Where minerals are absent at low pH, it is often the result of their high solubility in water at these pH values. As a consequence, they move down through the soil dissolved in water (i.e. they are leached) and so are taken into groundwater well below the level of plant roots. In contrast, phosphorus, manganese and boron are locked up in insoluble iron and aluminium phosphates at around pH 7.5. Plants growing on these soils can suffer from deficiency diseases (for example, phosphorus deficiency, which shows as yellowing of the leaves). This is particularly true of plants that are not adapted to grow at high pH. Phosphorus is also unavailable at low pH; again it is locked up as insoluble aluminium phosphate.

Limestone-derived soils may have very high pH values — 8 or above — particularly where the soil is shallow and directly overlies carbonate rock. Very few acid grasslands have a pH as low as 4.5. However, low pH values are a feature of moorlands characterized by heather (*Calluna vulgaris*) and bracken (*Pteridium aquilinum*), and also of upland bogs characterized by rushes (*Juncus* spp.) and cotton sedge (*Eriophorum vaginatum*), where the pH can be below 5.5.

The reasons why plants show preferences for acid or alkaline soils are not always simple, i.e. attributable solely to nutrient availability. For instance, some minerals can be toxic. Thus calcicoles grown on acid soils suffer from aluminium and other toxicities that cause stunting of their root systems. This stunting reduces nutrient uptake as a secondary effect.

Salinity, desiccation and biotic interactions on sea-shores

Tidal movements ensure that sea-shore habitats are, if not covered by seawater for part of each day, at least subject to spray-borne salt and wind. So, even well above the level of high tides, sea-shore organisms need to be more tolerant of salt than most terrestrial organisms. However, salinity (the concentration of salts dissolved in water) is not the only factor affecting sea-shore species. Seaweeds and shelled animals like limpets and barnacles are adapted to living in a highly saline marine environment, but they suffer desiccation if they are out of water for too long between high tides. In addition, where the shore is in an exposed location (i.e. a high energy environment), sea-shore organisms need to be able to tolerate severe abrasion caused by both wave action and pebble movements.

These factors affect not only which species are present in the intertidal zone, but also their exact location on the shore, i.e. their zonation^G. The pattern of zonation that results from these environmental factors is then further modified as the result of biotic interactions between organisms, such as seaweeds and the molluscs that graze on them. So, patterns of zonation vary and interpretation of what is found can be rather complicated.

Figure 4.10 shows the zonation pattern of seaweeds and grazing animals on two rocky shores.

Question 4.4 Describe the differences between the occurrence and zonation of channelled wrack (*Pelvetia canaliculata*) and limpets (*Patella* spp.) on these two shores. ◀

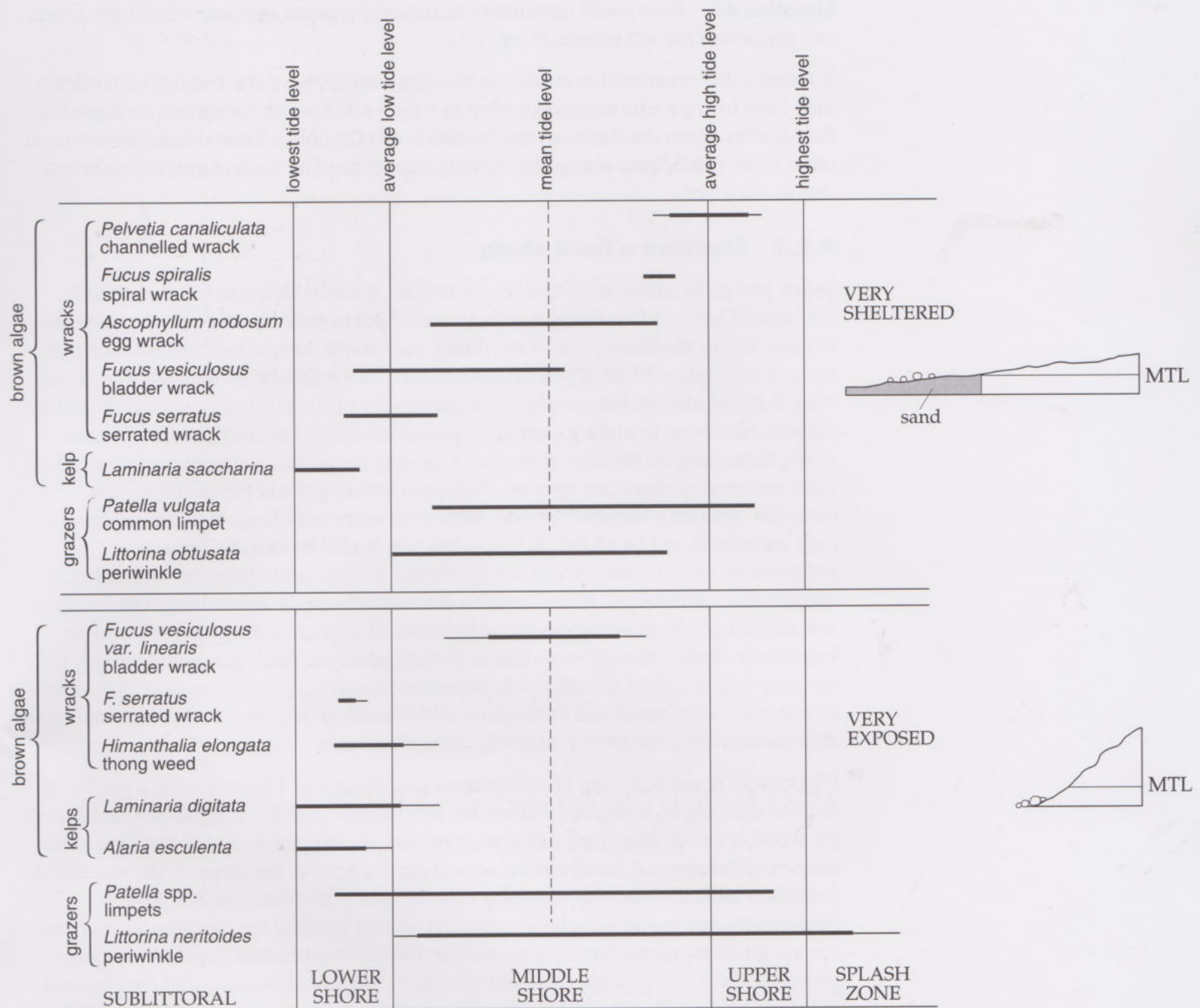


Figure 4.10 Generalized zonation pattern of seaweeds (algae) and grazing animals on two rocky shores. (MTL = mean tide level)

Channelled wrack is a species that is very resistant to desiccation and can recover even when it loses 96% of its water content. However, it has a relatively low growth rate and so it may not be able to become established on shores where young plants would be damaged by wave action. Other factors affecting the growth and zonation of channelled wrack include its inability to tolerate low light levels (for instance, when covered by the more profuse growth of other seaweeds) and its susceptibility to disease when submerged in water for long periods of time.

Limpet distribution on the lower reaches of sheltered shores is linked to the heavy growth of brown seaweeds, such as wracks and kelps. When juvenile limpets (spat) are released into seawater they must settle on rock surfaces to develop further. If no rock surfaces are exposed there is no place for them to settle.

Question 4.5 How could the salinity of the environment *increase* when tides recede and organisms are left exposed? ◀

It follows that organisms occurring in the spray zone, above the level of high tides, may have to cope with an environment in which salinity can sometimes be higher than further down the shore, where the tide is out for only a limited time. However, at other times rainfall can reduce the salinity experienced by these organisms to below that of seawater.

4.2.3 Starting a field study

When you go to somewhere that is new to you, and of which you know very little, you should have a plan of action to help you to put together a scientific description of the site. Fairly obviously, one of the things you would do is to note the location and make a description of what you are able to see. This might be in the form of a sketch map. It might also include producing a species list of the plants and animals found in the site. However, to make a complete species list might take days, weeks or even years, depending on the size of the site. One way round this is to make a list of the most common or dominant species. With time you might add the names of less common, yet fairly frequent species. With even more searching, the species that are only occasional can be added. In doing this you would be making subjective judgements, i.e. judgements that are qualitative and do not require the collection of quantitative information. But subjective assessments can be misleading. For instance, you might be influenced by the 'attractiveness' of a species, the prominence of its location or simply through bias, that is finding what you think you *ought* to find. In the long run, as a field scientist, it is important to collect quantitative data that ensure an unbiased assessment and description of the location. It is also often important to be able to describe a site over a relatively short time-scale.

Figure 4.11 is a sketch map of a deciduous woodland site. Figure 4.12 is a profile of the site showing its topography along the line marked A to B. You can see that part of the wood is on top of a small hill, part is on a slope and part is on a flatter area at the bottom of the slope. A small stream runs along the base of the slope. Is the woodland a uniform habitat? One way of testing this out is to walk along the line A to B (technically this would be called a transect), noting whether the tree species that you see are the same throughout. Let us assume for the moment that they fall into three groups (Figure 4.13). Ashes and oaks (Figure 4.14b and d) dominate on the flat land at the top of the hill, oaks and birches (Figure 4.14c) dominate on the slope, and alders (Figure 4.14a) and birches along the narrow stream-side strip and on the flat area beyond. If these three areas really are distinct, it would be safe to survey each of them as if they were separate habitats. But to ensure that there was no bias in your data collection, your three areas should be sampled randomly. If the changes between the top and bottom of the hill are gradual, in other words if it is not easy to see where one habitat ends and another begins, then you need to sample at intervals right along the transect.

Woodland is usually a multi-layered (i.e. three-dimensional) habitat, with tall trees forming the tree layer^G, small trees forming the understorey^G and low-growing shrubs and herbaceous plants forming the ground layer^G. Figure 4.13 is a profile of these three layers along the transect A–B. The key tells you that, at the top of the hill, the understorey includes holly (*Ilex aquifolium*); however, this species does not occur on the slope or at the foot of the hill. Similarly, the ground layer at the top of the hill is thick with bluebells (*Hyacinthoides non-scripta*) in spring but these do not occur on the slope or at the foot of the hill. Near the stream there are big clumps of one of the

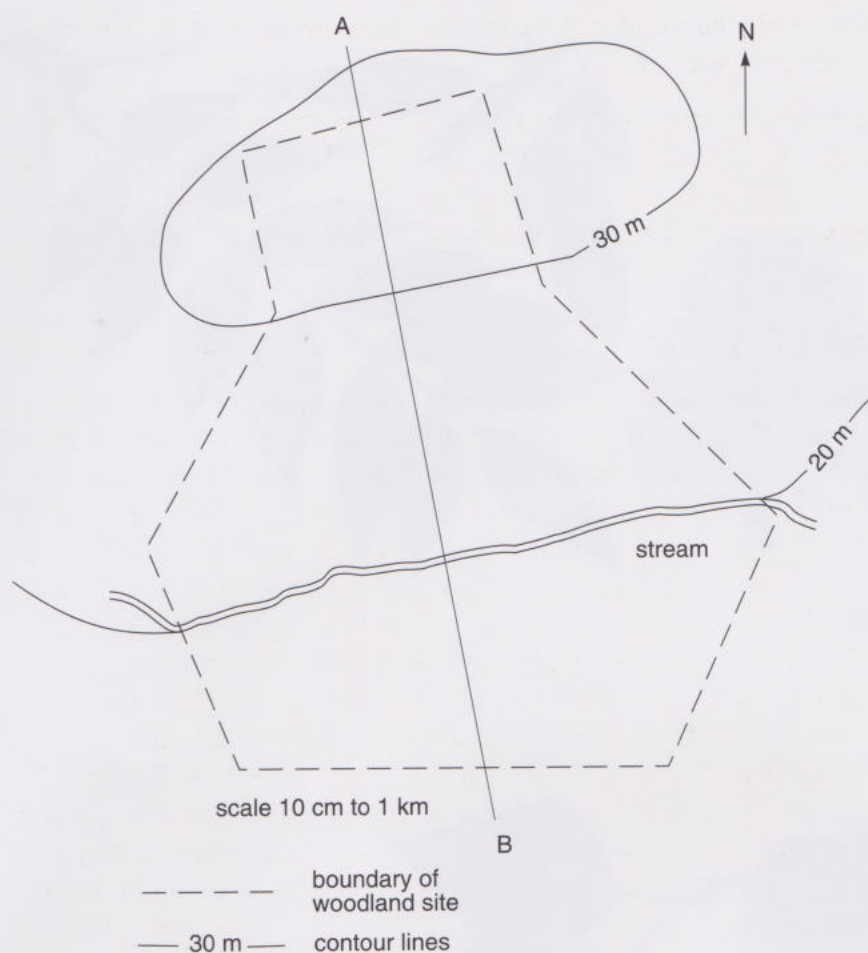


Figure 4.11 Sketch map of a hypothetical deciduous woodland that runs down from a hill top to a flat area at the bottom of a slope.

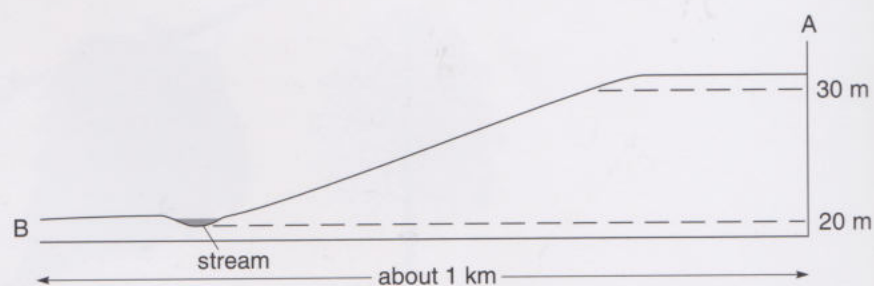
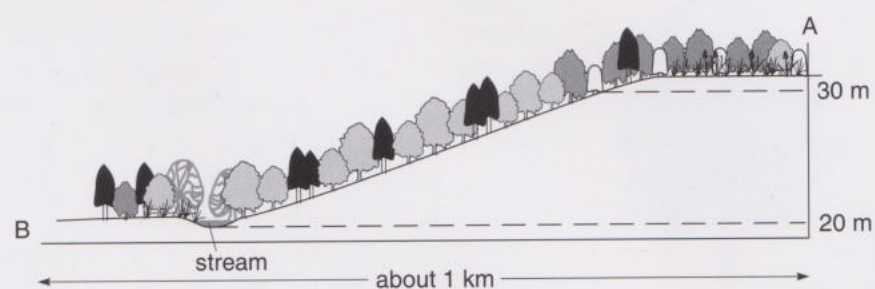


Figure 4.12 A topographical profile along transect A-B through the hypothetical deciduous woodland in Figure 4.11.



- | | |
|-----------------------------------|--|
| oak (<i>Quercus</i> sp.) | bluebells (<i>Hyacinthoides non-scripta</i>) |
| ash (<i>Fraxinus excelsior</i>) | sedges (<i>Carex</i> sp.) |
| birch (<i>Betula pendula</i>) | |
| alder (<i>Alnus glutinosa</i>) | |
| holly (<i>Ilex aquifolium</i>) | |

Figure 4.13 Profile showing the tree, understorey and ground layers along transect A-B through the hypothetical deciduous woodland in Figure 4.11.



Figure 4.14 Leaf shapes and outlines of mature trees of (a) alder (*Alnus glutinosa*), (b) ash (*Fraxinus excelsior*), (c) silver birch (*Betula pendula*) and (d) sessile oak (*Quercus petraea*).

tall grass-like sedges (*Carex* spp.). From plant distributions in either the tree or the ground layer it is reasonable to propose that the three sections of the profile are different. If you think you can differentiate between areas, it is reasonable to think they may be different habitats and therefore they should be described individually. You can, however, provide better evidence of this by collecting quantitative data.

Question 4.6 Figure 4.15 is a profile of the vegetation in a spruce plantation. (a) Describe how the vegetation (other than the trees themselves) differs from the profile of deciduous woodland (Figure 4.13). (b) How could you account for the differences? ◀

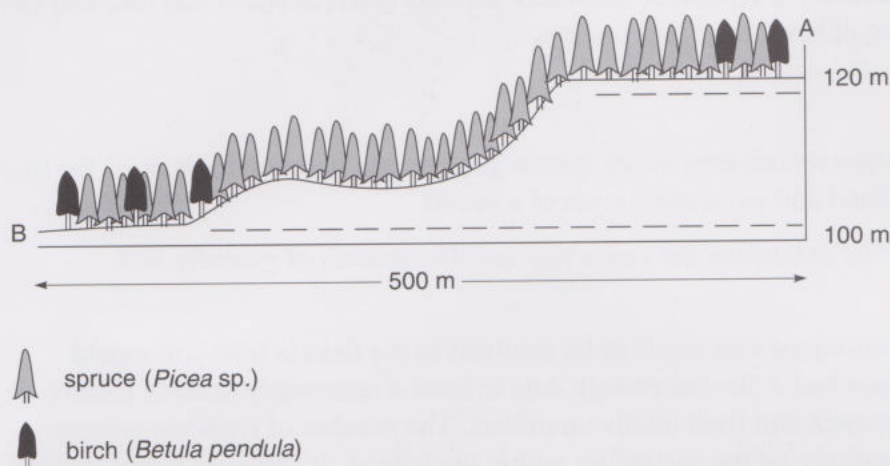


Figure 4.15 Profile of the vegetation along a transect through a plantation of spruce (*Picea*).

Most habitats are three-dimensional to some extent, although the scale may not be as noticeable as it is in deciduous woodland. Garden lawns, for instance, may well have a close ground-hugging moss layer under the grass and this can become dominant, to the annoyance of the lawn-keeper, if the soil becomes compacted and hence poorly aerated.

4.2.4 Collecting quantitative data

How can you collect quantitative data that summarize the nature of a habitat when it is three-dimensional? How, in fact, do you collect quantitative data?

Multi-storied habitats where the components have a different scale are usually recorded storey by storey, but using much the same methodology. There are two standard ways of collecting data quantitatively. The first involves recording species present within a standard area such as 10 cm × 10 cm, 0.5 m × 0.5 m, 1 m × 1 m, or 10 m × 10 m. It's all a matter of scale: 1 m × 1 m or smaller would be used as the unit for recording species in the ground layer whereas 10 m × 10 m would be used to record trees. Any device used for this purpose, whether it is one or more tape measures or a wooden frame, is called a quadrat^G. Within individual quadrats it is often possible to record the actual numbers of each species, or their presence or absence within each of a number of quadrats to show their frequency. Alternatively, the area occupied by each species can be recorded to give cover measurements.

A modification of the square quadrat that is more accurate — so particularly useful in habitats that have long lists of species — is the point quadrat^G. In such environments, sorting out which species are present (and especially the area they occupy) in even a 0.5 m × 0.5 m quadrat can be very laborious. Point quadrats are, in essence, minimal area quadrats. Each species touched by a sampling pin as it passes through the vegetation before reaching the soil surface is recorded as being present at that point.

An alternative to area counts is time counts, i.e. recording the number of organisms seen in a set period of time (e.g. 1 or 5 minutes) (assuming that diligence does not alter!). It may not surprise you that this method is generally used to record numbers of organisms that move around. Thus, flying insects can be collected in a sweep net which is moved back and forward as the collector walks onward for a given time. Ground-living, but mobile, insects such as beetles or slugs can be collected in submerged traps that the researcher visits once or twice a day. Woodland mammals, such as mice, voles, squirrels and stoats, are collected in traps that close upon the animal when they enter it (bedding and food *must* be available while the mammal remains trapped because recording should never be allowed to affect the size of the population). Numbers of relatively sedentary animals (such as snails and limpets) can be counted using either area or time counts.

Question 4.7

- (a) How would you record data about species present and their numbers in (i) the tree layer of a woodland and (ii) a moss layer of a lawn?
- (b) How could you determine the frequency and distribution of greenfly in a garden? ◀

One of the last questions that needs to be resolved in the field is how you would know whether you had collected enough data to have a reasonably reliable picture of the organisms present and their relative numbers. The number of replicate samples made depends entirely on the variability within the habitat. If there are many species you need more replicates to get reliable results, i.e. results that tell you which organisms are truly 'dominant', 'frequent' or 'occasional'.

4.2.5 The Residential School

On the fieldwork day at your Residential School you will learn some relatively simple methods that are widely used by ecologists to quantify data in terrestrial habitats. Having studied this preparatory material, you should be in a good position to contribute to discussions about the factors that may be important in the locations you are investigating and to reach conclusions about the data you collect in the field.

You should now read Section 3 of the workbook for Activity C, 'Investigating the environment', for the locations that you will visit during the Residential School.

4.3 Summary of Section 4

The learning outcomes for Part I of Activity D are listed in the workbook.

Sections 4.1.1 and 4.1.2 summarize the biology that you are expected to know for Activity D before commencing *Practising Science*, whereas Section 4.1.3 comprises Part I of Activity D itself.

The following paragraphs summarize the ecology covered in Section 4.2.

Ecology is the study of the interactions between organisms and their environment (including other organisms). An understanding of ecology is important to inform environmental decision-making.

Soil pH influences the availability of mineral nutrients to plants and hence the distribution of different plant species. Some species may be classified as either calcicoles or calcifuges.

Variation in salinity, exposure to desiccation and biotic interactions (e.g. grazing) influence the zonation of seaweeds and animals on rocky shores.

Field studies usually involve the collection of quantitative data in as objective a way as possible. Transects are often taken across field sites where some sort of environmental gradient is suspected. Quadrats at various scales, including point quadrats, are used to estimate frequency or the proportion of area occupied by different plant species. Time counts may be used to sample organisms that move.

Now that you have completed Section 4 you should be able to:

- define ecology;
- explain some of the reasons why the study of ecology is important;
- describe some of the ways in which soil pH influences the distribution of organisms;
- describe some of the reasons why seaweeds and some animals display zonation on rocky shores;
- describe in general terms how transects, quadrats, point quadrats and time counts can be used objectively to collect quantitative data about field sites.

5

Chemistry: analysis, reactions and structures

5.1 Introduction

Chemistry is about the structures of materials their properties and reactions. The chemistry input into the Residential School is associated mainly with Activity B 'Analysing our environment', and Activity E (Part II) 'The activities of metals'. The chemistry in Activity B is concerned with how we analyse our environment. How do we know what pollutants are present? How do we know how much of the substance there is? What of an unknown substance — how do we unravel its structure? We spend a lot of time with colourful reactions and spectra, yet carry out careful analysis at the part per million level! Activity E involves more familiar chemical processes. We look at transformations - acids, bubbling solutions and hot metals. We answer questions such as why did the Bronze Age come before the Iron Age. The work requires careful observation and measurements, but the results are very significant to the way we use different metals for different purposes.

This preparatory material is not an introduction to chemistry from scratch. As you will see, we expect you to have some familiarity with *basic* chemical concepts already. Its aims are to reinforce these chemical concepts and also the maths skills that underpin the laboratory work you will undertake in chemistry. At the same time we will introduce a few chemical concepts that you may not have met before.

This preparatory material uses the workbooks for Activities B and E to introduce the various chemical concepts described, so you will need to have them to hand. You should try to answer the questions posed in the workbooks, writing your answers in pencil in the spaces provided. You will be able to discuss your answers with your tutor at the Residential School. Don't worry if you are unsure of any answers — they will be much more obvious when you carry out the experiments for yourself.

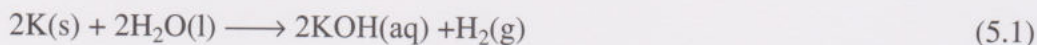
Start by reading Section 1 from Part II of the Activity E workbook 'The laboratory work: the activities of metals'.

5.2 The reactivity of metals with hydrochloric acid solution

We assume you know that the alkali metals from Group I of the Periodic Table (reproduced here in Appendix 3), such as sodium and potassium, are all highly reactive. The metals react quickly with water, in some cases explosively, to form a metal hydroxide and hydrogen gas. We can write a word equation for this that shows how the reactants are converted into products:

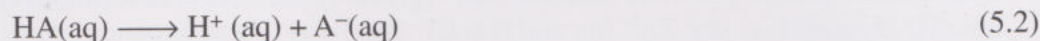
solid potassium plus liquid water gives potassium hydroxide solution plus hydrogen gas

However, as chemists, we use chemical formulae to generate a balanced equation, where there is the same number of atoms of each element on either side of the equation:

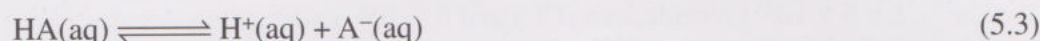


Question 5.1 The metal calcium (Ca) also reacts with water to give calcium hydroxide $\text{Ca}(\text{OH})_2$ and hydrogen gas. Write a balanced chemical equation for this reaction. ◀

In the first part of the experiment we examine the reaction of some metals with hydrochloric acid (HCl). Remember, an acid is a substance that contains hydrogen and yields hydrogen ions when it dissolves in water. For example, we can write a chemical equation for the dissociation of the hypothetical acid, HA, in water.



This is an equilibrium^G process, which is designated by the special arrow $\rightleftharpoons$:



The arrow in each direction highlights the fact that the reaction is in a state of balance – at the same time as reactant molecules are converted into H^+ and A^- ions in the forward reaction, $\longrightarrow$, the product H^+ and A^- ions are converted into reactant molecules in the back reaction, $\longleftarrow$. At equilibrium, the rate at which product is formed in the forward direction matches the rate at which product is lost in the back reaction, so there is no overall change in the concentrations. Thus although we have a state of balance, at the molecular level there is still great activity as reactants are converted into products and vice versa.

The idea of equilibrium also reinforces the idea that reactions don't always proceed to completion, that is, reactions don't always involve reactants being completely converted into products. Some reactions reach a state of balance where a mixture of reactants and products is formed or even very little product. This position of equilibrium is reflected in the equilibrium constant^G, K , which for the above reaction is expressed as:

$$K = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \quad (5.4)$$

Remember, the square brackets round a formula mean 'concentration of', so $[\text{HA}]$ refers to the concentration of the acid HA.

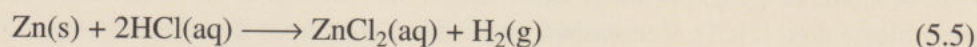
If the equilibrium lies to the side of products, that is, the final mixture consists of mainly H^+ and A^- ions, then K will be large (greater than 1). However, if the equilibrium lies to the side of reactants, that is, the final mixture consists of mainly HA molecules, then K will be small (less than 1).

For the above reaction, if HA is completely dissociated^G into H^+ and A^- ions, so there is essentially no molecular HA left in solution, we say that HA is a strong acid^G. If the acid exists mainly as molecular HA in aqueous solution with only a small fraction being converted into H^+ and A^- ions, we say the acid is a weak acid^G.

The HCl you will use in Activity E is a strong acid: thus a 6 mol litre⁻¹ solution of HCl contains 6 mol litre⁻¹ of hydrogen ions. Remember that the term mol litre⁻¹ is a measure of concentration it tells us how many moles of substance are contained in a litre of solution. If you need reminding about the mole^G, have a look at Box 5.1.

Box 5.1 The mole

We start with the following chemical equation which is already correctly balanced:



Quite simply, it means that *one* atom of zinc reacts with *two* molecules of hydrochloric acid to yield *one* formula unit^G (given by the chemical formula) of zinc(II) chloride (i.e. one Zn^{2+} ion and two Cl^- ions) and *one* hydrogen molecule. In a similar way we could write that 6×10^{23} atoms of zinc react with *two* times 6×10^{23} molecules of hydrochloric acid to yield 6×10^{23} zinc ions (Zn^{2+}), $2 \times 6 \times 10^{23}$ chloride ions (Cl^-) and 6×10^{23} hydrogen molecules. We didn't choose the figure 6×10^{23} out of thin air. It is Avogadro's number — the number of particles of a substance in a mole of that substance. For example, one mole of hydrogen molecules contains 6×10^{23} particles (molecules in this case) of hydrogen and one mole of zinc also contains 6×10^{23} particles (atoms in this case). So an easier way of interpreting the equation above is that *one* mole of zinc atoms reacts with *two* moles of hydrochloric acid to yield *one* mole of zinc(II) chloride and *one* mole of hydrogen molecules.

How does one mole, or 6×10^{23} atoms, of zinc, relate to a mass of zinc? Put simply, it is the relative atomic mass of zinc in grams, i.e. the relative atomic mass of zinc is 65.4, so one mole of zinc has a mass of 65.4 grams. Table A3.1 in Appendix 3 lists the values of relative atomic mass that you need to study this preparatory material.

For compounds we need to add up the relative atomic masses of the atoms in the chemical formula. The molar mass^G of a compound i.e. mass of one mole, is then the sum of the relative atomic masses in the chemical formula, in grams (the unit is grams per mole, or g mol^{-1}). Similarly, for an element, the molar mass is the relative atomic mass in grams.

- Given that the relative atomic mass of chlorine (Cl) is 35.5, what is the molar mass of zinc(II) chloride?
- The chemical formula contains one zinc ion and two chloride ions, therefore the relative molecular mass is $65.4 + 2 \times 35.5 = 135.4$. The molar mass of zinc(II) chloride is therefore 135.4 g mol^{-1} .

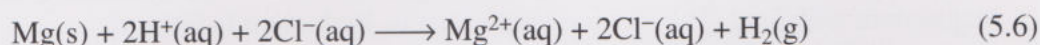
The balanced Equation 5.5 shows that 1 mole of zinc reacts with 2 moles of hydrochloric acid to form 1 mole of zinc(II) chloride and 1 mole of hydrogen gas. Then, 65.4 g (1 mole) of zinc reacts with $2 \times (1.0 + 35.5) = 73 \text{ g}$ (2 moles) of hydrochloric acid to form 135.4 g (1 mole) of zinc(II) chloride and $2 \times 1.0 = 2 \text{ g}$ (1 mole) of hydrogen gas. Note the conservation of mass: 137.4 g of reactants form 137.4 g of products.

The acidity of a solution depends on the concentration of hydrogen ions (H^+). This is usually measured using a scale known as the pH scale, which you may use in Activity C if you are going to Sussex University. If the hydrogen ion concentration is written as $1.0 \times 10^{-n} \text{ mol litre}^{-1}$, then the pH is simply equal to n . For example, if a hydrogen ion concentration is $1.0 \times 10^{-2} \text{ mol litre}^{-1}$, then the value of n is 2 and the corresponding pH is also 2. A pH value is not necessarily an integer: for example, a pH of 3.7 corresponds to a hydrogen ion concentration of $1.0 \times 10^{-3.7} \text{ mol litre}^{-1}$; $10^{-3.7}$ has a value of 2.0×10^{-4} , which you can prove for yourself by using your

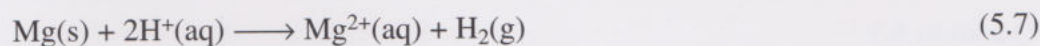
calculator — key in 10 then push the x^y button followed by -3.7 (you may have to use the change sign button to get the minus). Thus a pH of 3.7 corresponds to a hydrogen ion concentration of $2.0 \times 10^{-4} \text{ mol litre}^{-1}$.

Now read Section 2 of Part II of the Activity E workbook to just before the start of Section 2.1. Don't worry too much if you don't understand all the experimental details, they will make more sense in the laboratory!

In this part of the experiment you will observe that magnesium reacts with hydrochloric acid to form an aqueous solution of magnesium chloride and hydrogen is given off as bubbles of gas. Magnesium chloride is an ionic compound so we can write the equation:



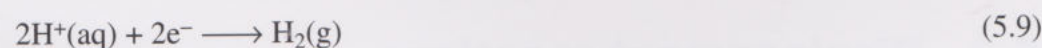
Notice that the chloride ion, $\text{Cl}^-(\text{aq})$, remains unchanged during this reaction, we say it is a 'spectator ion', and so we can simplify the equation by eliminating it from both sides to show only the species involved in the reaction:



(Remember that although this equation shows only positive ions these are balanced by the spectator ions, which don't appear.)

Question 5.2 Write down an atomic/molecular description of Equation 5.7 in words. ◀

To help us to understand what is going on in the reaction between magnesium and hydrochloric acid in aqueous solution we shall separate the reaction into two *half reactions*, which highlight what happens to each of the reactants:



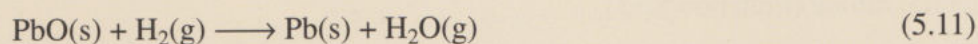
In Equation 5.8 neutral magnesium atoms lose two electrons, each denoted by e^- , to form positive ions Mg^{2+} and, in Equation 5.9, hydrogen ions (H^+) gain electrons to form hydrogen molecules. You should confirm for yourself that if we add Equations 5.8 and 5.9 together we end up with Equation 5.7. Notice that in a balanced equation the numbers of electrons gained and lost are the same. In fact substances gaining or losing electrons provide us with a way of defining whether species are oxidized or reduced as described in Box 5.2.

Box 5.2 Oxidation and reduction

Oxidation^G was originally defined as the addition of oxygen to something. For example, magnesium burns in oxygen to give magnesium oxide:



Oxygen has been added to magnesium, so by this definition magnesium has been oxidized. The converse, reduction^G, was defined as the removal of oxygen from a compound: for example, if lead oxide is heated in a stream of hydrogen gas it is reduced to metallic lead:



Note that the hydrogen has gained oxygen and hence been oxidized in this reaction — oxidation and reduction always complement each other.

These original definitions are fine when oxygen is added or removed from something — but there are a great many reactions that involve oxidation and reduction in which gain or loss of oxygen is not evident. In fact, we examined such a case in Equation 5.6 when magnesium lost electrons to form magnesium ions, Mg^{2+} . This loss of electrons constitutes a new definition of oxidation, i.e. Oxidation Is Loss of electrons (the mnemonic OIL might help!)

If something loses electrons during a reaction then something else must have gained them — Equation 5.9 shows it is the hydrogen ions in the reaction of magnesium with acid. This is a reduction, that is, Reduction Is Gain of electrons. (The mnemonic here is RIG.) Because any oxidation is counterbalanced by a reduction (and vice versa) we can use the collective mnemonic OIL – RIG.

Question 5.3

(a) Write down a balanced chemical equation (including states) for the reaction of calcium metal with aqueous hydrochloric acid to give calcium chloride (CaCl_2) in solution and hydrogen gas.

(b) Write this equation in full ionic form. Identify spectator ions and remove them to give the simplest ionic equation. ◀

You should now read the rest of Section 2 of Part II of the Activity E workbook. Questions 1, 4 and 5 involve balancing equations and oxidation and reduction. Your tutor will discuss these questions with you at Residential School, but you should have a first attempt at these questions now.

5.3 Displacement reactions

In chemistry we rarely look at the reactivity of a substance in isolation. It is generally more instructive to see reactivities in a context, i.e. within families of substances, and this provided the basis of the Periodic Table. For example, the properties of the element sodium are more clearly understood when compared with those of the other elements in its group and also with the other elements in its period. Just about any metal can be placed within a reactivity sequence but we will focus on only those metals that are in general use and/or ubiquitous in chemistry simply because they are more common. Such a series of reactivity is particularly important because it tells us how easy it is to extract a metal from its ores.

In Part 3 of the experiment we create such a reactivity series^G by studying the displacement reactions^G of metals in metal salt solutions. Have a look at the three stages shown in Figure 5.1. This shows what happens when a piece of copper wire is dropped into a solution of silver(I) nitrate (AgNO_3). The copper wire initially blackens and, after a while, is coated with beautiful needle-like crystals of a silvery-white substance. Also the solution, which was colourless, turns light blue. The blue colour arises from the formation of Cu^{2+} ions from metallic copper (Equation 5.12) and the silvery-white metal is silver, formed from the silver ions (Ag^+) produced from silver(I) nitrate (Equation 5.13).

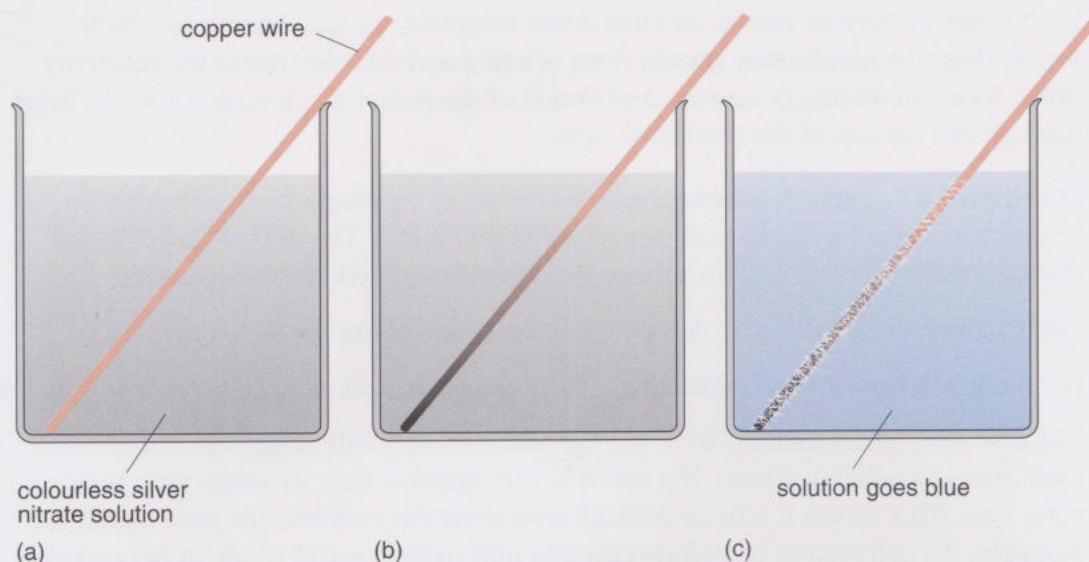
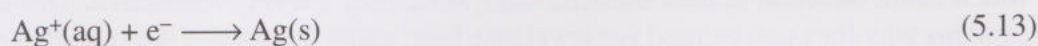
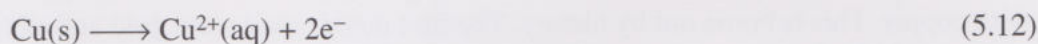
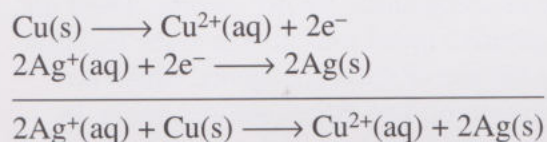


Figure 5.1 The displacement reaction of metallic copper in silver nitrate solution: (a), (b) and (c) represent various stages of the reaction.



Equations 5.12 and 5.13 are half-equations. What is the full ionic equation for this reaction?

To get the full equation, we add the two half-equations together so that the number of electrons lost by the copper atoms equals the number of electrons gained by the silver ions. Two electrons are exchanged, so we need to add Equation 5.12 and two times Equation 5.13, then cancel out any species that appears on both sides of the arrow:



In this reaction we say that the copper metal has displaced the silver ions from the solution to give metallic silver, and in doing so the copper has been converted into copper ions.

Which metal is oxidized and which reduced in this reaction?

The copper atom loses electrons so is oxidized (OIL), and the silver ions gain electrons so are reduced (RIG).

One way of defining the reactivity of metals is based on their willingness to give up electrons and become oxidized — the more readily the metal gives up its electrons the greater its reactivity. Earlier, we discussed how sodium metal readily reacts with water and is oxidized to form sodium ions (Na^{+}), so we can say that sodium is very reactive. Now, copper and silver are completely unreactive to water so they are clearly less reactive than sodium, but how reactive is copper in relation to silver?

The displacement reaction in Figure 5.1 shows copper forming ions at the expense of silver ions. We can put this another way by saying that copper has a greater tendency than silver to form ions, thus copper is more reactive than silver. The reactivity order for the three metals is:

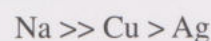


Table 5.1 Reactions of metals with metal ions.

Metal salt solution	tin	lead	zinc
tin nitrate	–	×	✓
lead nitrate	✓	–	✓
zinc nitrate	×	×	–

By combining various metals and metal salt solutions, we can determine which metals displace which other metals from solution and thus determine the reactivity order. You will do this in Section 3 of Part II of the Activity E laboratory work. Read through this section of the workbook now.

Question 5.4 Table 5.1 summarizes the results of dipping a piece of metal (tin, lead or zinc) into the nitrate solution of a *different* metal. The ticks ✓ indicate that reaction was observed, and the crosses × indicate no reaction was observed.

(a) Use these observations to deduce the order of reactivity for the three metals.

(b) Write a balanced ionic equation for the reaction of lead nitrate solution with tin. ◀

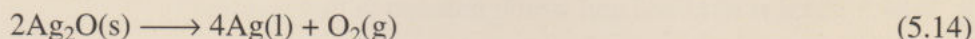
The reactivity series enables us to predict which metals will be easy to extract from their ores and which difficult. If a metal is very reactive then its atoms will readily form ions. This means it will be difficult to convert the ions into the pure metal; for example, the conversion of sodium chloride into sodium metal is not an easy process. However, a less reactive metal will be more easily formed from its ions, as is the case with copper. This is borne out by history. The first metals used were gold and silver, which could be found in their metallic state, since they are very unreactive. This also explains why they can be used for jewellery. Next came the Bronze Age and then the Iron Age, because copper, tin and iron are fairly unreactive and thus could be extracted from their ores using simple processes. The more reactive metals, such as sodium and potassium, were isolated only relatively recently, because they are difficult to extract. The various extraction processes are discussed in Box 5.3.

Box 5.3 The extraction of metals

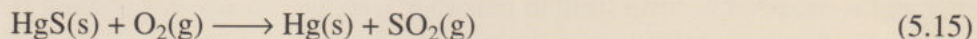
Only the least reactive metals, such as gold and silver, are found in their metallic state; all of the others occur as compounds, known as ores. There are, broadly, three ways of reducing the positive metal ions in ores to the metallic state. All involve heat, some involve the use of reducing agents, which are capable of readily being oxidized themselves, and others use electricity. In practice, the method used depends upon the reactivity of the metal concerned.

Heating the ore

This works only for very unreactive metals. For example, silver can be obtained from its oxide (Ag_2O) by heating it to a temperature greater than 160°C :

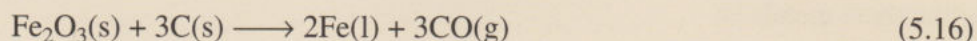


Sometimes the ore needs to be first converted into its metal oxide by gentle heating in air. Subsequent more vigorous heating then decomposes the metal oxide to the metal and oxygen. For example, mercury (Hg) a metal of low reactivity occurs as the sulfide (HgS) in the ore cinnabar, which when heated in air gives the metal:



Addition of reducing agents

Reducing agents such as carbon (in the form of coke) can be used to extract metals with moderate reactivity, such as tin, iron and zinc. For example, heating iron oxide with coke in a blast furnace gives iron:



Electrolysis

The more reactive a metal is, the more forcing conditions are needed to obtain the metal. With the most reactive metals, such as those found in Groups I and II of the Periodic Table (magnesium, calcium, sodium and potassium), electricity is required to force enough electrons into the ionic compound to ensure that the positive ions accept the electrons and become reduced to neutral metal atoms:



The whole medium has to be molten because electricity will not conduct through ionic solids — the ions need to be free to move. This process is known as electrolysis.

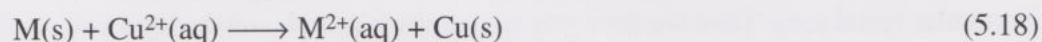
This is how aluminium is obtained. Bauxite (Al_2O_3) is mixed with cryolite (sodium aluminium fluoride) to make it easier to melt. Electricity is then passed through the molten mixture and the molten aluminium collected in pots.

5.4 Enthalpy change for a displacement reaction

All chemical reactions involve either a giving out or a taking in of energy. We see this when we burn a fuel in air and energy is given out in the form of heat from the reaction of the fuel with oxygen. The source of this energy is the difference in potential energy^G between the reactants and products. All reactions involve breaking and making of chemical bonds or the transfer of electrons. Thus the energy that is taken in or released during a reaction is the result of a difference between the energy stored in the bonds of the reactants and the energy stored in the bonds of the products.

The energy that is given out or taken in during a reaction in the form of heat (at constant temperature and pressure) is known as the enthalpy change^G for a reaction, and is given the symbol ΔH . When heat energy is released, ΔH is negative and the reaction is said to be exothermic^G. When heat energy is taken in, ΔH is positive and the reaction is said to be endothermic^G.

The displacement reactions that are seen to occur in Section 3 of the Residential School experiment are all exothermic reactions. Thus, heat energy is released and in this series the amount of heat energy released reflects how readily one metal displaces another — the more heat energy released the greater the difference in reactivity between the metals. In Section 4 of the experiment we measure the amount of heat energy released when a series of metals, M (magnesium, aluminium, iron, tin and zinc) displace copper from a solution of copper sulfate:



Because we use the same metal salt, copper sulfate, in all cases, the amount of heat energy released will give us a numerical measure of the relative reactivities of magnesium, aluminium, iron, tin and zinc.

Now read Section 4 of the workbook for Activity E until just before Section 4.1. Again the experimental procedure may not be obvious without seeing the apparatus for yourself, but it should give you a good idea of the chemistry involved.

Section 4.1 involves the calculation of the enthalpy change for the reactions. By convention we always calculate the energy released or taken in when the amounts

match those in the balanced equation in molar terms. Thus, for Equation 5.18, we would calculate the heat energy released when one mole of the metal M reacts with one mole of copper ions. Thus, the last section involves scaling up the measured heat change, which is related to a certain mass of copper ions being displaced, to estimate the heat change that would occur if a mole of copper ions were displaced.

In fact, the enthalpy change refers to the energy in the form of heat that is transferred from the system to the surroundings, at constant pressure and temperature. In this case the system can be considered as the solution of copper sulfate and the metal, whereas the surroundings are everything else, including the polystyrene cup. In this experiment we estimate the highest temperature that the solution achieves, and thus calculate how much energy was evolved to achieve this temperature rise by this mass of solution. This then enables us to calculate the amount of energy in the form of heat that is transferred from the system to the surroundings when the solution cools to the original temperature.

Have a look at Section 4.1 of the Activity E workbook now.

Question 5.5 The equation used to calculate the amount of heat q from a measured temperature rise is given in Section 4.1 of workbook for Activity E. What would be the heat required to cause a temperature rise of 20°C in 50 g of copper sulfate solution? (Use the value for specific heat of solution given in the workbook.) ◀

Questions 11, 12, 13 and 14 in Part II of the Activity E workbook involve enthalpy and concentration. Your tutor will discuss these questions with you at Residential School where you will use your own value of q determined from the activity. You might like to practise these questions now, using the value for q you calculated in Question 5.5.

5.5 Activity B 'Analysing our environment'

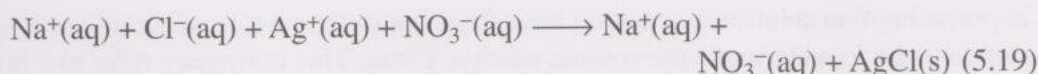
Activity B uses much of the chemistry covered in Activity E, but introduces two new concepts: a chemical test for an unknown substance, and the calculation of concentrations.

5.5.1 A chemical test for an unknown substance

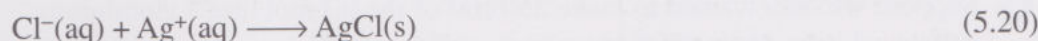
Suppose we had a salt consisting of an unknown metal ion (positive) and an unknown negative ion, such as sulfate, chloride or carbonate. Chemists have created a series of chemical tests that will identify which metal ion is present. Similarly, another series of chemical tests has been developed for the negative ion, so that the overall identity of the material may be discovered. In the first part of Activity B, we ask you to carry out a series of reactions on known metal ions and thus identify some key tests for particular metal ions. Then we give you an unknown sample, of the kind you might get in any analytical laboratory, and ask you to determine its identity.

Most of these tests involve precipitation reactions. Essentially two soluble salts are mixed, for example, aqueous sodium chloride (containing Na^+ and Cl^- ions) and aqueous silver nitrate (containing Ag^+ and NO_3^- ions).

Whereas sodium chloride, silver nitrate and sodium nitrate are soluble in water, silver chloride is not and thus forms a white precipitate when the two solutions are mixed. Note that the state symbol (aq) indicates the soluble species and the state symbol (s) denotes the formation of a precipitate.



- Rewrite Equation 5.19 without the spectator ions.
- The spectator ions are $\text{Na}^+(\text{aq})$ and $\text{NO}_3^-(\text{aq})$; their removal from Equation (5.19) leaves



Some of these precipitated salts are coloured, and this provides another clue to the identity of our unknown sample.

Read through Part A1 of the Activity B workbook to see how such reactions will be used to identify unknown metal ions. Try to complete Table A2 in Part A1, listing the positive and negative ions of the metal salts in the table before you attend the Residential School, where your tutor will discuss the answers with you.

Notice that you are given the formula and charge of the nitrate ion so you can work out the charge on the positive ion(s) in the compound, because the number of negative charges on the negative ions must balance the number of positive charges on the metal ion. For example, chloride ion is Cl^- , and the chemical formula of magnesium chloride is MgCl_2 . The formula indicates that one magnesium ion is associated with two chloride ions, so overall there are two negative charges (2Cl^-). This is balanced by the charge on the magnesium ion, which must therefore carry a 2+ charge — Mg^{2+} .

5.5.2 Calculating concentrations

In Part A2 of Activity B you will dilute a standard solution^G (one of known concentration) of aluminium sulfate to give a range of solutions whose concentrations you will be able to calculate on the basis of the dilution factors. These solutions are then used to calibrate an instrument known as a colorimeter, which you will use to measure the concentration of aluminium in an unknown water sample. You then assess whether the aluminium concentration in this water sample is within acceptable concentration limits set for drinking water. Remember, chemists usually talk about concentrations in terms of mol litre^{-1} . This is equivalent to the amount of the substance in moles contained in 1.0 litre of solution.

Suppose the standard solution is a $0.100 \text{ mol litre}^{-1}$ solution of aluminium sulfate (in fact it is much more dilute than this). This concentration is equivalent to that obtained by adding sufficient water to 0.100 mole of aluminium sulfate to make up a litre of solution.

- What is the molar mass of aluminium sulfate, $\text{Al}_2(\text{SO}_4)_3$? (You will need to use the values for relative atomic masses given in Table A3.1 in Appendix 3.)
- The relative molecular mass of $\text{Al}_2(\text{SO}_4)_3$ is $\{2 \text{ Al } (= 2 \times 27.0) + 3 \text{ SO}_4 (= 3 \times (32.1 + 4 \times 16.0))\} = 342.3$, therefore the molar mass is 342.3 g mol^{-1} .
- How many grams of aluminium sulfate will one litre of $0.100 \text{ mol litre}^{-1}$ aluminium sulfate contain?
- One litre of $0.100 \text{ mol litre}^{-1}$ aluminium sulfate contains $0.100 \times 342.3 \text{ g} = 34.2 \text{ g}$ of aluminium sulfate.

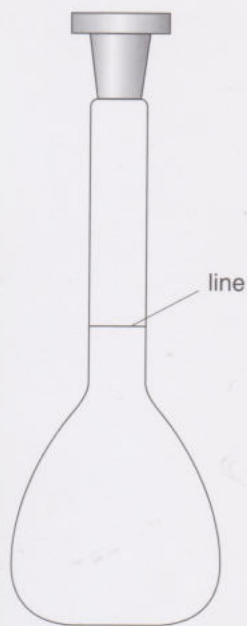


Figure 5.2 A volumetric flask.

In preparing this solution you would have to weigh out the 34.2 g of aluminium sulfate and add water until the volume reaches a litre. This is why we refer to a litre of solution each time: you would not add a litre of water because 34.2 g of aluminium sulfate will occupy some volume. The precise volume, 1000 cm^3 , is achieved by using a special piece of glassware called a volumetric flask, pictured in Figure 5.2. The line is marked at 1000 cm^3 .

- Suppose we only wanted to make 250 cm^3 of the $0.1 \text{ mol litre}^{-1}$ aluminium sulfate solution. How much aluminium sulfate would we put in our 250 cm^3 volumetric flask and make up to 250 cm^3 of solution?
- A volume of 250 cm^3 is 0.25 litre ($250 \text{ cm}^3/1000 \text{ cm}^3$), so we need only use a quarter of 34.2 g, which is 8.55 g.

Let's now carry out the calculation the other way round, starting with the mass of the salt and working out its concentration.

- A solution of aluminium sulfate is made by taking 1.0 g of aluminium sulfate ($\text{Al}_2(\text{SO}_4)_3$) and adding sufficient water to make 200 cm^3 of final solution. What is its concentration in mol litre^{-1} ?
- The molar mass of aluminium sulfate is 342.3 g mol^{-1} . Thus 1.0 g of aluminium sulfate is equivalent to $1.0 \text{ g}/342.3 \text{ g mol}^{-1} = 2.9 \times 10^{-3} \text{ mol}$. This is contained in 200 cm^3 . The amount that would be contained in 1 cm^3 of this solution is $2.9 \times 10^{-3} \text{ mol}/200 \text{ cm}^3$, and the amount that would therefore be contained in a litre of this solution is $(2.9 \times 10^{-3} \text{ mol}/200 \text{ cm}^3) \times 1000 \text{ cm}^3 = 1.5 \times 10^{-2} \text{ mol}$. The concentration is therefore $1.5 \times 10^{-2} \text{ mol litre}^{-1}$. *Notice that the final answer is given to the same number of significant figures (two in this case) as the initial 1.0 g amount of aluminium sulfate.*
- If we take 25 cm^3 of this solution and add sufficient water to make a final volume of 100 cm^3 , what is the final concentration?
- The simplest way to do this type of calculation is to determine by what factor the solution has been diluted. We start off with 25 cm^3 and end up with 100 cm^3 so we have increased the volume by a factor of $100/25 = 4$ and thus diluted the solution by a factor of 4. Thus the final concentration is $1.5 \times 10^{-2} \text{ mol litre}^{-1}/4 = 3.8 \times 10^{-3} \text{ mol litre}^{-1}$.

Activity 5.1

In this activity you will practise carrying out the dilution calculations and graph plotting needed for Part A2 of Activity B. You will use this graph to estimate the aluminium concentration of an unknown water sample, given the absorbance, and thus decide whether the water is suitable for drinking, based on the acceptable levels.

- (a) Read through Part A2 of the Activity B workbook. Using the hypothetical data in Table 5.2, calculate the concentration of aluminium in mol litre^{-1} for each solution (the concentration of the standard solution is $8.0 \times 10^{-5} \text{ mol litre}^{-1}$).
- (b) On the graph paper in Figure 5.3, plot a graph of absorbance against the concentration that you calculated in Table 5.2, and draw the best-fit straight line through the origin.
- (c) An unknown water sample has an absorbance of 0.75. Using your graph, read off the concentration of aluminium in mol litre^{-1} in the unknown solution that corresponds to this absorbance and thus calculate the concentration of aluminium in p.p.m. as discussed in Task A4 of the Activity B workbook. ◀

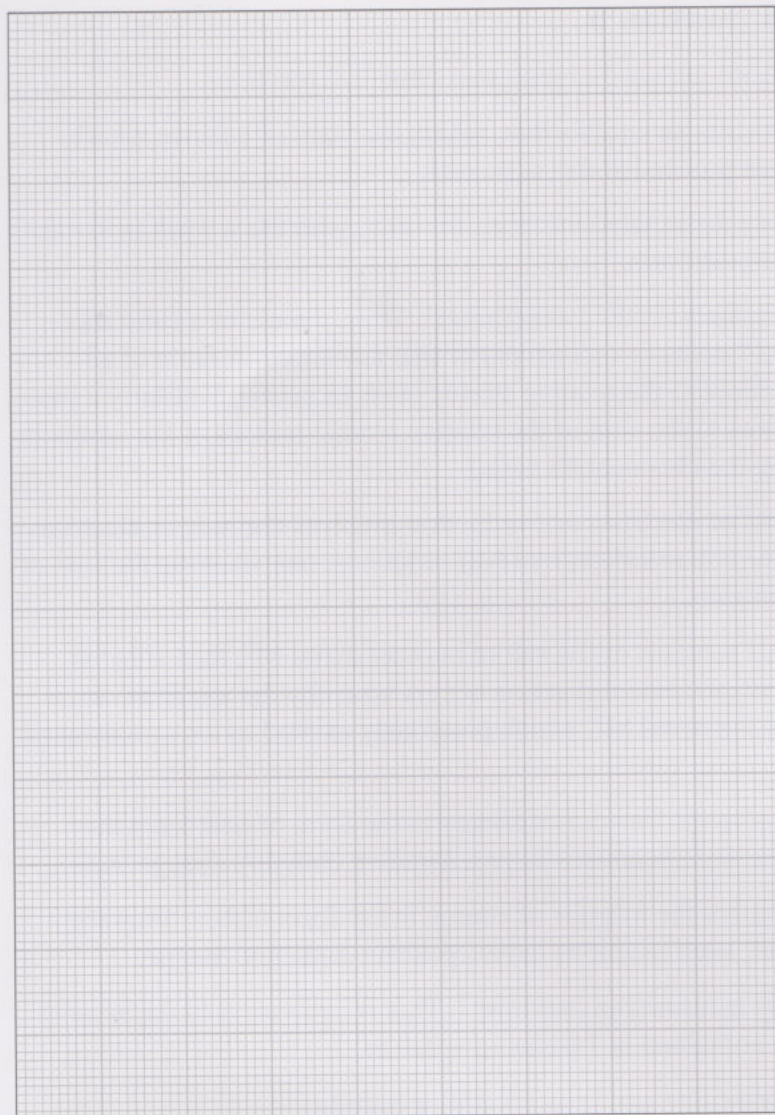


Figure 5.3 Graph paper for use with Activity 5.1.

Table 5.2 The dependence of absorbance on concentration.

Solution number	Number of 10 cm ³ portions of standard solution	Concentration of aluminium/mol litre ⁻¹	Absorbance
Distilled water	0		0
1	2		0.40
2	3		0.60
3	4		0.90
4	5		1.15

5.6 Summary of Section 5

Balanced chemical equations are used to represent chemical reactions.

The symbol $\rightleftharpoons$ represents an equilibrium, which is a state of balance where forward and back reactions proceed at the same rate.

A strong acid is completely dissociated into ions in water, whereas a weak acid exists predominantly in its molecular form.

A mole of something contains 6×10^{23} 'somethings'. Thus, a mole of a substance contains 6×10^{23} particles (atoms or molecules) of that substance. A mole of an element has a mass equivalent to the relative atomic mass in grams. Similarly a mole of a compound is equivalent to the sum of the relative atomic masses of the atoms in the formula unit (given by the chemical formula).

Oxidation involves the loss of electrons and reduction involves the gain of electrons (OIL RIG). We can divide such reactions into half-reactions, one showing the loss of electrons and the other the gain.

If a metal M^1 is placed in a solution of the salt of another metal M^2 , then if M^1 is more reactive than M^2 , a displacement reaction occurs whereby the metal M^1 is oxidized to form ions and the metal M^2 , which previously existed as ions, is reduced to the pure metal.

Exothermic reactions involve the transfer of energy in the form of heat from the system to the surroundings; such reactions have a negative enthalpy change, ΔH .

Endothermic reactions involve the transfer of energy in the form of heat from the surroundings to the system; such reactions have a positive enthalpy change.

Concentrations are usually measured in mol litre^{-1} . This is equivalent to the amount of the substance in moles contained in 1.0 litre of solution.

Now that you have completed Section 5 you should be able to:

- write balanced chemical equations for the reaction of hydrochloric acid with metals;
- interpret chemical equations as descriptions of reactions either at the atomic level (atoms, ions and/or molecules), or at the macro level (moles of reactants and products);
- explain oxidation in terms of loss of electrons, and reduction as gain of electrons;
- deduce the relative activities of metals from their ability to undergo displacement reactions;
- describe the nature of displacement reactions, that they are exothermic and that the magnitude of ΔH for displacement (of copper from copper sulfate by a range of metals) gives an indication of the relative activities of metals via the numerical magnitude of the enthalpies determined for these displacement reactions;
- identify metal ions in solution using simple reactions that form precipitates;
- calculate the concentrations of solutions in mol litre^{-1} and parts per million from amounts expressed as grams of solute dissolved in a volume of solution expressed in cm^3 .

If you have experienced difficulties with the mathematics in this section, you may find the section 'Maths Help' in the *SGSG* useful (page 302 of *SGSG* lists the areas covered).

Research project: aluminium – exploring its impact

6

6.1 The purpose of this project

You will already have noted that this project has a chemistry investigative component — a study of some metals, including aluminium, with the aim of placing them into a reactivity sequence. Background material for this part of Activity E is provided in Section 5 and includes preparatory material required for the chemistry in Activity B as well.

In this section we are concerned solely with the research and group work aspect of this activity. In particular, we want to make sure you are aware of the rationale behind Activity E and the type of work you will be undertaking. We also want to develop some of the abstracting skills that you will use.

You should now read Part I of the Activity E workbook, 'Research project: aluminium – exploring its impact', which explains what you will be doing.

Have another look at the list of skills you should develop during these activities, listed at the end of Section 1 of Part I. At first these skills may not appear to be very scientific but they are, nevertheless, essential skills that scientists need to develop. For example, the main method by which scientists tell each other of their research is through publication of papers in scientific journals. It is an important scientific skill to be able to glean the key points that are relevant to your research from a paper, and thus write a brief summary. This enables you to keep all the information in a more manageable form. If you need to expand on anything, you know where the ideas came from.

A less obvious scientific skill is communication to a lay audience, through a range of different media. We have just discussed how important it is for scientists to be able to communicate their results to each other and it is becoming increasingly important for scientists to explain what they do to a wider audience. The general public is usually their paymaster, either through the taxes paid to government or as consumers, so it is important that the public understands what scientists do. Even though scientists may be experts in their own areas of science, they will still need scientists from other areas to communicate with them in lay terms – this may be as part of studies, in the hospital, in the field or in the laboratory. This activity enables you to become an 'expert' in a particular field and to communicate your knowledge to others. In so doing, you will appreciate some of the generalizations that have to be made and how difficult it is to avoid jargon!

You may have noticed that many people have very different and firmly held views about topics of current scientific interest and public debate, such as genetically modified organisms (GMOs) and BSE/CJD. If we think about these different views as a spectrum, they can range from pro-science at one end, through ambivalence in the middle, to complete opposition to scientific advances at the other end. Each of these positions can alter according to the science under consideration. So, for example, an individual who is opposed to developments in stem cell research could be enthusiastic about the search for extra-terrestrial life.

It is important to note that most scientific developments rarely gain wide levels of public discussion. When they do, it is often because they are deemed controversial with

arguments based on a range of evidence, scientific or otherwise. For example, ethical, moral and religious arguments can often be seen as challenging the scientific consensus on a range of issues, including those already mentioned above. The danger here, for scientists, is that current levels of scientific knowledge, which can inform these debates, can be lost in this contested and often emotionally charged environment. The fault lies at least in part with the scientists. They usually prefer to stay in their laboratories and get on with their research. When they do publish their findings, it is usually in specialized journals using language unfamiliar to non-specialists. To address these problems, all science students in UK universities are being taught how to communicate scientific ideas and information to both a specialist and a lay audience.

Much of this impetus for the need to improve communication was started by a report on the Public Understanding of Science produced by a Royal Society Committee chaired by Walter Bodmer in 1985. However, this is still a very live issue: for example, in 1999 there was a House of Lords report on the Public Understanding of Science, and its significance in the future will only increase as the need for informed public debate becomes more widespread. We hope that the scientists of the future will be able to communicate the details of their work, their interests and their passions to a wide audience and thus overcome some of the prejudices that the general public have about science.

Teamwork is another key skill of the scientist. Almost all science today consists of a team of scientists working towards a common goal. It may be a team of only two or three, or it can be a team of hundreds. In Activities A to D, you may work at some tasks in pairs, some in small groups and some alone. As you saw when you read Part I of the Activity E workbook, in this activity we specifically ask that you work as part of a team and you will need to communicate *within* the team. At the end of the week your group's work will be presented as a collection of posters — the result of an effort, through communication, by *all* the members of the team over a period of one and a half days.

6.2 The scope of this project

What sort of science are we asking you to consider in this project? Where is the information for this science going to come from? As you have discovered, the focus of the science is the metal aluminium and its compounds. The information will largely be supplied by us, in one form or another. However, there is no reason why you should not be on the look out for information on aluminium and its compounds and bring this with you. We don't want you to spend any time poring over journal articles, or trawl endlessly through the Internet: merely keep your eyes open for any general information – such as may be found in newspaper or magazine articles. Don't worry if you don't come across anything, because we will have plenty of sources at Residential School, such as:

- Journal articles
- Newspaper reports
- Video films
- The Internet
- The laboratory work associated with this project
- Your Group Tutor

Working in a team, each of you will engage in certain skills essential to the abstraction of information. To prepare you for the kind of thing you will be doing, we are including here, in Activity 6.1, some tasks based on a single article about acid rain. The article is now, at the time of writing, thirteen years old and therefore not fully up to date. This does not matter: we are using it purely because it is an article that still has relevance and is useful for you to practise with. Notice also that certain effects of aluminium do figure strongly in it.

Activity 6.1

Read the article 'Acid rain' (on pages 96 to 99) and answer the following questions:

- (a) (i) Find two errors in the chemistry related to sulfate and hydrogen ions in this article.
- (ii) The sulfate ion is derived via chemical reactions from sulfur dioxide: oxidation is said to occur. What definition of oxidation is being used here?
- (iii) Write balanced equations for the sequences:
 The conversion of $\text{SO}_2(\text{g})$ into $\text{SO}_3(\text{g})$ by reaction with oxygen.
 The conversion of $\text{SO}_3(\text{g})$ into $\text{H}_2\text{SO}_4(\text{aq})$ by reaction with water.
- (b) (i) Briefly describe the processes whereby SO_2 is transported to the ground (either as SO_2 or as H_2SO_4).
- (ii) What is the main source of SO_2 in Europe, and what steps are being taken to reduce this source?
- (iii) Name two other gases that give rise to acid rain, and two major sources of these.
- (iv) If rain falling over the Pennines has a pH of 3.5 and the pH of the water droplets in clouds from which this rain is falling is 'typically ten times more acidic', what is the pH of the water droplets in the clouds?
- (c) (i) Give one piece of evidence implicating post Industrial Revolution acid gas emissions in acid rain?
- (ii) Carbonic acid is a *weak* acid; sulfuric acid is a *strong* acid. What do the italicized terms mean in the context of acids?
- (d) Describe:
 - The link(s) between acid rain and fish depletion in rivers and lakes;
 - The link(s) between acid rain and depletion of forests. Why is this a more complicated issue? ◀

Activity 6.2

Provide a summary of the whole of the article 'Acid rain' in not more than 150 words. [Hint: to do this you will have to identify the key points that are made and make brief notes about each point. You may find that your first summary is too long, in which case you will need to summarise it further!] ◀

Activity 6.3

In not more than 400 words, summarize the discussion in the article 'Acid rain' concerning the relationship between atmospheric pollutants and damage to plants and animals. ◀

Reproduced from *New Scientist, Inside Science*, 5 November 1987, pp. 1-4.

New Scientist 5 November 1987

INSIDE SCIENCE

ACID RAIN

A decade ago, few people had heard of acid rain. Now, it is a major political issue. But what is acid rain? Where does it come from? And, more important, what can we do to minimise its effects?

Fred Pearce

newscientist
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Acid rain is not a recent discovery. A hundred years ago, Britain's first air pollution inspector, Robert Angus Smith, coined the phrase to describe the polluted rain in his home town of Manchester. He realised that the air in the city was not only filthy but also acidic, and that it was attacking vegetation, stone and iron.

The finding, and the phrase, were forgotten, however, until the 1960s. Then Scandinavian scientists began to link pollution blown across the sea from Britain and the European mainland with acidified lakes and streams and the disappearance of fish from those waters. At first, no one could explain why acid rain affected rivers when vast amounts of acids occurring naturally in soils did not. Still less had they established what killed the fish.

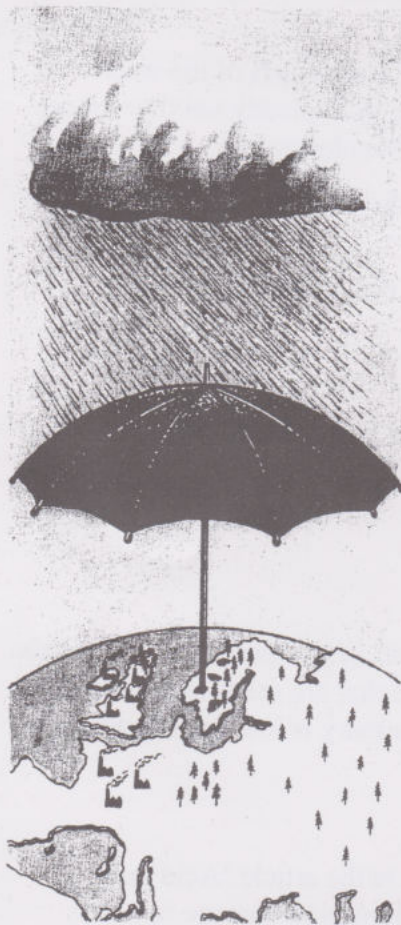
Rain is naturally slightly acidic – it reacts with carbon dioxide in the air to produce weak carbonic acid with a pH of about 5.6. In central Europe, rain is far more acid with an average pH of around 4.1. Even on the western fringes of the continent, such as Ireland and Portugal, pH averages 4.9. Rain from individual storms can have a pH of less than 3, and water droplets in fogs may be more acid still.

Acid rain has become an important political issue over the past 15 years, souring relations between the polluters and the polluted: between Britain and Norway, for instance, and the US and Canada. Air pollution can spread for thousands of kilometres across land and seas.

Europe, the most concentrated seat of air pollution in the world, still shows the worst signs of acid damage, but there is evidence of damage from Australia to Mexico to China. Acid rain is now a global phenomenon. But until recently, the

science of acid rain has been confused.

Now, the key debate about soil chemistry and the transfer of acid from rain to rivers is largely over. Research into the atmospheric chemistry behind the formation of acid rain is beginning to yield important results. But the processes behind the destruction of Europe's forests – by far the most serious effect so far attributed to acid rain – are, to put it mildly, unclear.



In the clouds A cocktail of chemicals

One of the chief culprits in acid rain is sulphur dioxide, which reaches the atmosphere by many routes. Sea spray, rotting vegetation, plankton and, in some places, volcanoes are important natural sources. Round the world, perhaps half the sulphur in the air comes from such sources and half from burning fossil fuels. Over Europe, the proportion from burning fuel is about 85 per cent.

Once sulphur dioxide (SO_2) reaches the atmosphere, it reacts with moisture in the air. First, it is oxidised, taking up atoms of oxygen to become a sulphate ion (SO_4). This then combines with hydrogen to form sulphuric acid (H_2SO_4).

Not all the sulphur dioxide, however, becomes sulphuric acid. The process of oxidation can happen in dry air as a "gas phase" reaction but it is usually slow. A large mass of sulphuric dioxide, in the plume from a power station chimney for instance, can travel hundreds of kilometres in stable air with little sign of conversion to acid. Over much of Europe, most of the sulphur dioxide falls to the ground unconverted.

When sulphur dioxide is incorporated into clouds, however, oxidation by "wet phase" reactions seems to happen much more quickly. Researchers tracking plumes from power stations across the cloud-covered Pennines, found most of the sulphur dioxide converted to acid within a couple of hours. Droplets in clouds are typically 10 times more acid than the rain that falls from those clouds.

There are grounds for believing that the whole atmosphere over Europe and the US is chemically much more reactive than a few decades ago, and that this

INSIDE SCIENCE

A history of assault in Britain

Britain has an instructive history of assault by acid rain. Ever since the chimney became commonplace, ejecting fumes of burning fuel into the outside air, towns have suffered from air pollution. Smoke has been the most obvious problem. But sulphur, too, has had a part to play. Coal contains, typically, between 1 and 3 per cent sulphur which disappears up the chimney as sulphur dioxide.

In December 1952, a pollution disaster hit London. A deadly "smog" – a cold, black and sulphurous stew of air – hung over the city for almost a week, trapped by a blanket of warm air. It was the worst of a long series of "peasouper" smogs that had hit London. It killed about 4000 people.

Smog irritated bronchial tubes, which became flooded with mucus. People choked to death or suffered heart attacks as they fought for breath.

At the time, doctors blamed the copious quantities of smoke and sulphur dioxide that accumulated in the smog. Scientists now believe that the formation of highly acid particles may have been important. One estimate puts the pH of London smog of 1952 at 1.6, rather more acid than lemon juice.

After 1952, a public outcry led to legislation banning the burning of smoky fuels in most towns and cities. This, coupled with the arrival of cheap North Sea gas, the growing use of electricity and a decision by the

government to build the next generation of power stations outside urban areas, ensured that towns became vastly cleaner places.

But the government took no specific steps to limit emissions of sulphur dioxide from power stations burning coal and oil. Britain's output soared. By the mid-1970s, chimneys up to 300 metres high were putting more than 5 million tonnes of sulphur dioxide a year into the air over Europe. Other European countries emitted similar amounts. The result was that towns were cleaner, but sulphur dioxide spread in increasing amounts to the remotest corners of the continent. As later emerged, rainfall everywhere became more and more acid. □

accelerates the acidification. In heavily polluted urban air, tiny particles of metals such as iron and manganese catalyse the reaction. Elsewhere, gases such as ozone and hydrogen peroxide appear to be critical. One unexpected finding is that ammonia, given off in large quantities from open slurry tanks holding wastes from factory farms, may be an important catalyst, at least at a local level.

There are now strict controls on the emission of sulphur dioxide from

power stations in the US and several European countries. There, the rising stars of air pollution are nitrogen oxides, which are producing acid rain on a scale approaching that of sulphur dioxide.

Nitrogen oxides are given off by power stations (again) and car exhausts, and include both nitrogen dioxide (NO_2) and nitric oxide (NO). In the air, nitric oxide quickly converts to nitrogen dioxide, which is itself oxidised and changed into nitric acid (HNO_3).

In the soil Chemicals on the move

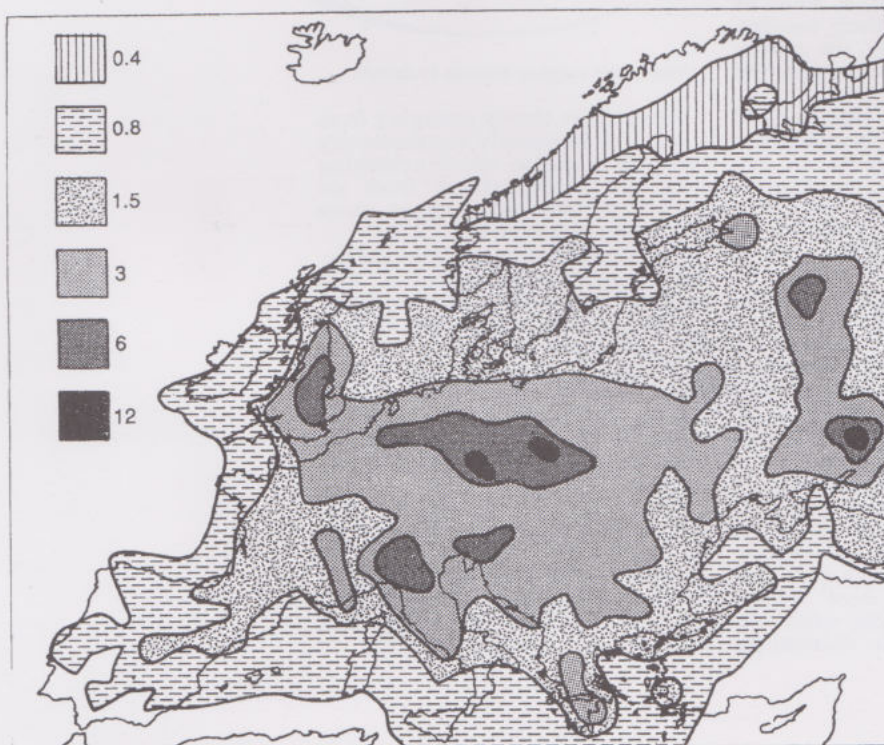
Many of the thin soils that cover granite or sandy rock in northern Europe, Canada and parts of the US have been acidic ever since they were formed 10 000 years ago at the end of the last ice age. Unable to neutralise acids that arise from natural processes, many of these soils contain the equivalent of thousands of years of deposition.

When the Scandinavians first claimed that pollution from Britain was causing acid lakes, critics pointed to this. It was areas with these soils that had many of the most acid lakes and rivers so, they asked, who needs acid to explain what is going on?

This theory weakened in the early 1980s when scientists reconstructed detailed histories of acid lakes. They looked in lake sediments for remains of tiny organisms called diatoms. These are very sensitive to acidity. It quickly emerged that lakes in Scandinavia, Scotland and Canada had become acid only after the height of the Industrial Revolution in the 19th century.

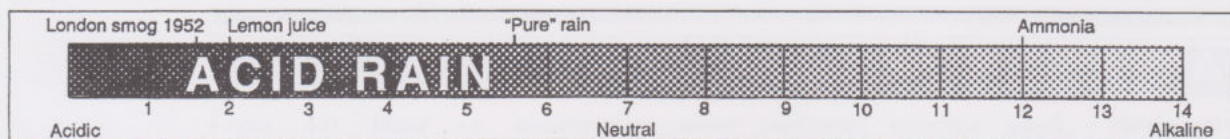
The clincher was a detailed study of soil chemistry. Natural acid in soils is dominated by carbonic and organic acids. But to transfer acidity to surface waters that drain the soil requires mobile negative ions to bind to the acid hydrogen ion. The carbonic and organic acids do not contain such ions.

The strong acids found in acid rain do, however. In particular, the negative sulphate ion in sulphuric acid is mobile within the soil and efficiently transfers acidity from



The total fallout of sulphur, in grams per square metre, over Europe in one year.
Nitrogen shows a similar pattern

INSIDE SCIENCE



The pH scale measures acidity and its opposite, alkalinity. It runs from 1 (very acid) to 14 (very alkaline); 7 is neutral. The scale is logarithmic so, for example, rain with pH of 4 is ten times as acid as rain with a pH of 5

soils to surface waters. The nitrate ion, too, would behave in this way if it were not normally taken up by plants first.

In the water The poisoning of fish

Fish, notably brown trout and salmon, have disappeared from thousands of lakes and several large rivers in southern Scandinavia since the 1950s. In 1900, anglers caught 30 000 kilograms of salmon in the seven main rivers in southern Norway. Since 1970, no salmon have been taken.

Many lochs in Scotland, especially in the Galloway Hills, are also fishless – as are hundreds more in Canada and parts of the eastern US. The fish died or failed to reproduce in acid waters, but only rarely is acidity the cause.

The deaths are usually due to poisoning by aluminium. All soils contain massive amounts of aluminium. Normally, though, it is in insoluble form, bound to the soil. But, just as the sulphate ion in acid rain transplants acid to rivers, so it can "unhook" aluminium from its complex compounds and wash it into streams. There it interferes with the operation of the fish's gills, so that they clog with mucus. This reduces the amount of oxygen that reaches the blood.

The mixture of aluminium and acid in lakes and rivers has a

profound effect on freshwater ecology. Acid lakes are usually crystal clear with luscious carpets of green algae and moss. All this green is deceptive. When algae and moss proliferate, they change the "metabolism" of the lake. They slow down the decomposition of the greenery, providing less energy for the life there. The result is a changed ecosystem, with fewer species.

Canadian scientists tried dosing a lake deliberately with acid over several years. The lake initially had a pH of 6.5. At a pH of around 6, shrimps and minnows were the first organisms to disappear. Trout eat minnows so young trout soon failed to appear. At 5.6, the external skeletons of crayfish softened and were soon infested with parasites. Their eggs were overrun by fungi. Soon there were no crayfish.

In the forest Trees under attack

The first signs of the decline of Europe's forests appeared in parts of the Alps, in the mid-1970s, when fir trees started to lose their needles. Then, in West Germany, the crowns of Norway spruce began to thin and needles turn brown.

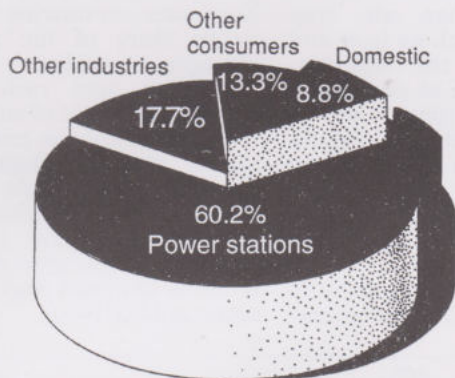
German foresters, in 1981, decided that the time had come to make their warnings public in newspapers and magazines. After 1984, the decline stabilised in Germany. In other countries, the forests continued to deteriorate, with deciduous trees increasingly affected.

In 1986, a European study classified about 29 per cent of trees in the Netherlands as moderately or severely defoliated. West Germany had 20 per cent in this category, and Czechoslovakia and Switzerland 16 per cent. A survey of selected sites in Britain recorded 29 per cent.

What is doing the damage? In Germany, scientists pointed early on to acid soils. Soils in parts of Sweden, West Germany and Britain are known to have become more acid in recent decades. Acid waters draining from the soils wash out nutrients and liberate aluminium,

which the roots of trees may take up. Without essential nutrients, such as magnesium and calcium, trees starve to death.

Sulphur dioxide also directly damages leaves and needles – it blocks the stomata on leaves, preventing photosynthesis. Ozone derived from vehicle exhausts reaches levels each summer that are toxic to some trees, especially in conjunction with sulphur dioxide. Shoots seem to develop at the expense of roots, photosynthesis is disrupted, and chemical processes in general upset.

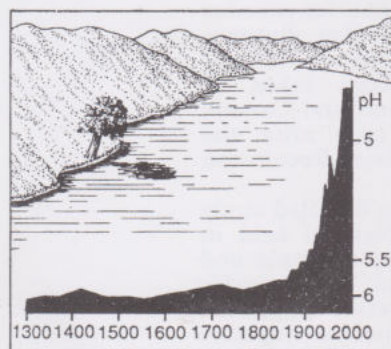


Who emits sulphur dioxide in Britain

The picture slowly emerging from a mass of frequently contradictory research is one in which acidifying soils and direct attack from air pollutants may make trees more vulnerable to assault.

Severe frosts may initiate decline. The suggestion, though, is that several air pollutants, including ozone and sulphur dioxide, reduce the frost-hardiness of plants. Air pollutants also encourage fungi and pests, such as the bark beetle, to grow. The ambrosia beetle is attracted by chemicals such as terpenes given off by trees under stress.

Ammonia, as we have seen, efficiently oxidises sulphur dioxide to create the sulphate ion. The resulting ammonium sulphate often forms on the surface of vegetation. Ammonia reduces frost-hardiness; ammonium sulphate, when it



How the acidity of a loch in Scotland soared with industrialisation

INSIDE SCIENCE

reaches the soil, creates both sulphuric and nitric acid.

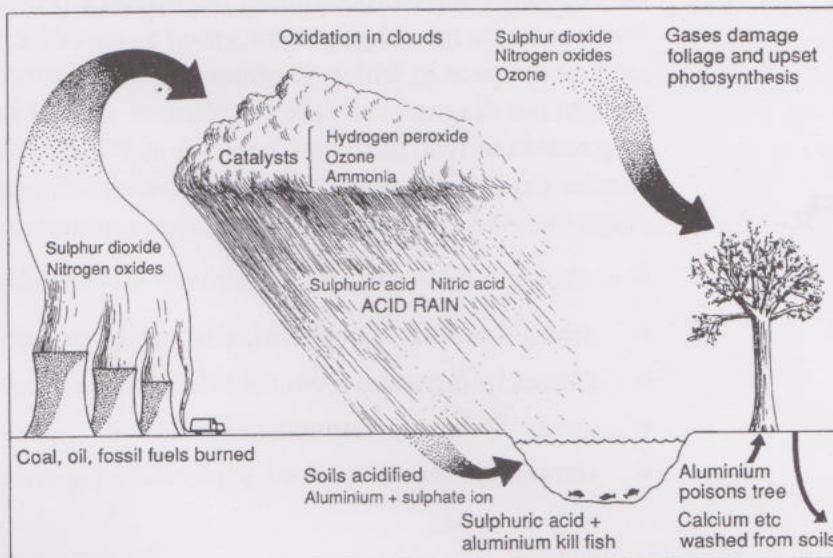
Some researchers believe that doses of nitrogen, in the form of nitric acid, nitrogen oxides or ammonium compounds, reach the soils in some parts of Europe in such quantities that they "saturate" the soils. We always assume that nitrogen does nothing but good to plants, but in excess it may stress trees by forcing them to grow when they are short of nutrients.

In the future Prospects for a cleanup

By the beginning of the next century, the output of sulphur dioxide and nitrogen oxides from industrialised nations will be much lower than today.

Most countries will install chemical plants to remove the sulphur from emissions from new power stations before they reach the atmosphere. A typical process involves large quantities of limestone (calcium carbonate) to remove sulphur dioxide and form gypsum (calcium sulphate). Careful design of combustion processes inside plants will reduce the output of nitrogen oxides from power stations.

The technology exists, too, to reduce pollution from the exhausts of vehicles. Catalytic converters,



The chain from pollutants to acidified lakes, disappearing fish and dying trees

which contain platinum catalysts, bolted onto exhaust systems remove hydrocarbons and nitrogen oxides. They are fitted as standard on cars in many countries, notably the US and Japan, and reduce the formation of both acid rain and ozone. The use of cars, however, is rising so rapidly that even stringent controls may do little more than maintain the status quo.

Even worse is that, if all air pollution were halted tomorrow, many of the effects outlined here

would remain for decades. Soils would retain their man-made acidity and reservoirs of sulphur. Acid and aluminium would continue to poison lakes and streams. A British attempt to model acidity in lochs in Scotland predicts that even a halving of acid fallout in the hills by the year 2000 would only maintain the acidity of water at their current levels.

The only way to ameliorate the effects of acidity in the short term is to combine a clean-up of the air with dumping vast quantities of limestone in soils and waters and adding fertilisers and nutrients to forest soils. One leading German scientist estimates that such a programme would cost his country £15 000 million.

It is a pessimistic picture, but at least research is beginning to unravel the complicated chemistry of acid rain. Once we understand it, we can address the problem. □

The bad side of ozone

Ozone is an enigma. In the upper atmosphere it is a "good thing". It shields the Earth from ultraviolet radiation. Close to the ground ozone is a hazard. It damages plants and many materials from rubber to textiles; it hastens the formation of acid rain, and may trigger asthma attacks and bronchitis.

Ozone is formed in sunlight by photochemical reactions between nitrogen oxides and traces of hydrocarbons in the air. Motor vehicles produce both – over Britain, around two-thirds of the ozone is generated by vehicle exhausts. Power stations are among the other sources of nitrogen oxides. Hydrocarbons come from everything from industrial solvents to the methane from ruminating cattle and leaking North-Sea gas.

Not surprisingly, the amount of background ozone close to the ground has roughly doubled over Europe in the past three decades.

There is a debate about how best to reduce the formation of ozone.

The amount of the fastest-reacting hydrocarbons, such as alkanes and alkenes emitted by vehicle exhausts, probably accounts for the peak concentrations of ozone in the summer. But there are also slow-acting hydrocarbons, such as methane, in the atmosphere which may take years to react with nitrogen oxide and create ozone. They are so common that ozone may best be stemmed by controlling nitrogen oxides.

Ozone is but one of the chemicals produced by reactions between pollutants in sunlight in industrial areas and which accelerate the formation of acid. One consequence of this increasingly reactive chemical soup in the atmosphere is the formation of heat haze, which is usually an aerosol of sulphates and nitrates. Such a soup, near Los Angeles, recently produced a fog with a pH of 1.7. □

FURTHER READING

Acid Rain by Fred Pearce (Penguin) is a popular account of the causes of acid rain, the damage it has done, and the remedies which need to be taken.

Ecological Effects of Deposited Sulphur and Nitrogen Compounds (The Royal Society) is the collected papers from a meeting at the Royal Society, and provides one of the best overviews of scientific knowledge in this complex subject. The text should be accessible to those with the equivalent of A-level biology or chemistry.

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6.3 What to do next

As you know, the culmination of the research project is the presentation of your findings to the rest of your tutor group as part of a poster. You will have only a small amount of space in which to summarize your work and you will find that a well thought out diagram can take the place of a lot of words. You will find it useful preparation to read through Chapter 3 in *SGSG* 'Working with diagrams'. This chapter explores a variety of ways in which information can be imparted through diagrams. As you will see, different types of diagram are used for different purposes.

Now that you have completed Section 6 you should be able to:

- read a scientific article critically and for understanding;
- extract information from the article that is relevant to a given question;
- use key points to summarize the article;
- summarize the discussions explored in the article.

Answers and comments for in-text activities

Activity 3.1 Calibrating a diffraction grating

Task 1

(a) (i) The wavelength λ is something that we know, since we are using a standard wavelength source.

(ii) The unknown quantity is d , the grating spacing.

(iii) The angle of diffraction θ_n is measured in the experiment. The order of diffraction n is also measured in the experiment (since we count the number of orders of diffraction from the straight through position).

(b) The two measured quantities are θ_n and n . Equation 3.4 does not describe a straight line relationship between these two quantities, because the equation is in terms of the *sine* of the angle of diffraction. The general equation of a straight line graph passing through the origin is $y = kx$. Equation 3.4 is

$$\sin \theta_n = \frac{n\lambda}{d}$$

This can be rearranged to give

$$\sin \theta_n = \left(\frac{\lambda}{d} \right) n$$

so there is a straight line relationship between $\sin \theta_n$ and n . So by taking the sine of the angle of diffraction, we should be able to obtain a straight line relationship.

(c) If $\sin \theta_n$ is plotted against n then we would expect the graph to be a straight line through the origin with a gradient of (λ/d) . So we could find d by measuring the gradient of the graph, and using the known value of λ ,

$$d = \frac{\lambda}{\text{gradient}}$$

Task 2

The values of $\sin \theta_n$ are shown in Table 3.4.

Table 3.4 A completed version of Table 3.1.

n	$\theta_n / ^\circ$	$\sin \theta_n$
0	0.0	0.000
1	12.1	0.210
2	25.1	0.424
3	39.2	0.632
4	57.5	0.843

Task 3

(a) Figure 3.21 shows a graph of $\sin \theta_n$ against n .

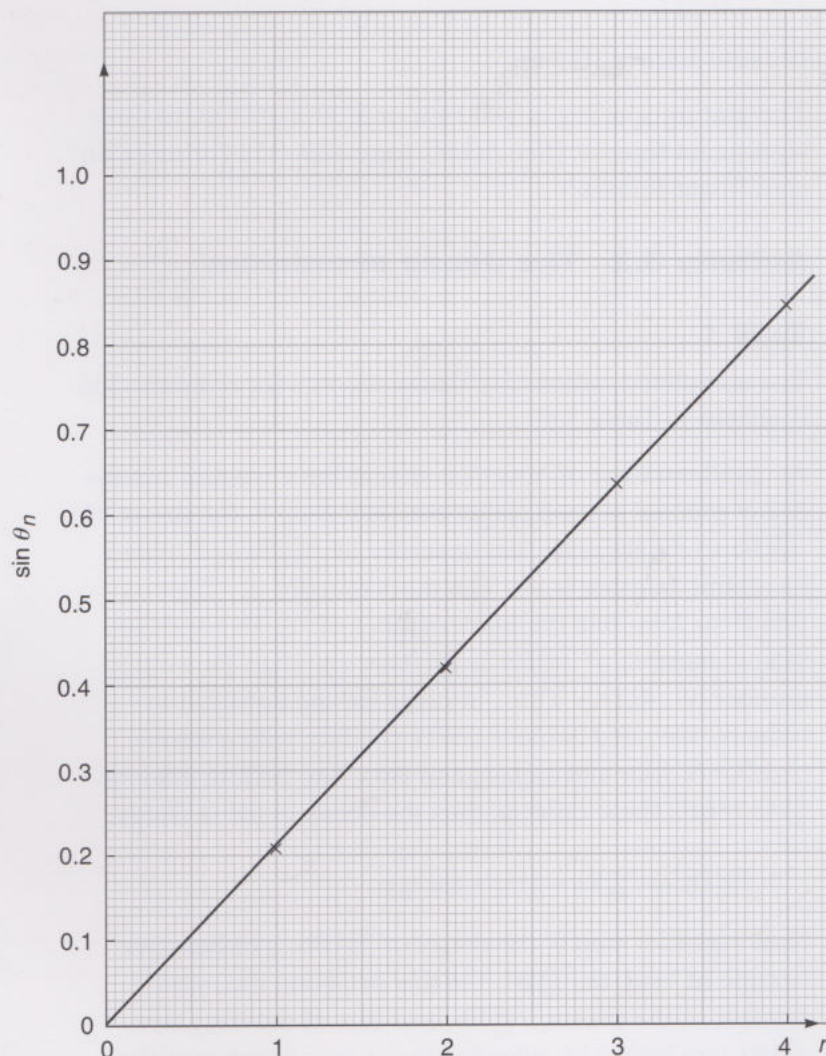


Figure 3.21 A graph of $\sin \theta_n$ against n (answer to Task 3a of Activity 3.1).

(b) The gradient is 'rise over run'. For the best fit line shown here,

$$\text{gradient} = \frac{\text{rise}}{\text{run}} = \frac{0.741 - 0.109}{3.5 - 0.5} = \frac{0.632}{3} = 0.211$$

Note that there is no unit associated with this gradient because there is no unit associated with either $\sin \theta_n$ or n .

(c) The grating spacing is

$$d = \frac{\lambda}{\text{gradient}}$$

so,

$$d = \frac{632.8 \text{ nm}}{0.211} = 2999 \text{ nm} = 2999 \times 10^{-9} \text{ m} = 2.999 \times 10^{-6} \text{ m}$$

because $1 \text{ nm} = 10^{-9} \text{ m}$.

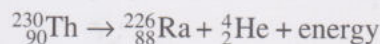
The grating spacing is 3000 nm or $3.00 \times 10^{-6} \text{ m}$, to three significant figures. (See *SGSG* page 349–352 for more on significant figures.)

Activity 3.2 The decay of uranium

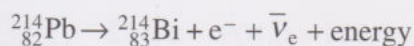
Task 1

The completed radioactive decay scheme for uranium-238 is shown in Table 3.5.

As an illustration of how you should have worked out the decay process at each stage, consider Step 5, the decay of $^{230}_{90}\text{Th}$ to $^{226}_{88}\text{Ra}$. The mass number of $^{230}_{90}\text{Th}$ is 230 and its atomic number is 90. The mass number of $^{226}_{88}\text{Ra}$ is 226 (four less than the mass number of $^{230}_{90}\text{Th}$) and the atomic number of $^{226}_{88}\text{Ra}$ is 88 (two less than the atomic number of $^{230}_{90}\text{Th}$). Thus, in order for the equation to balance, we need another decay product with a mass number of four and a charge of two. The helium nucleus ^4_2He meets these criteria, thus the decay is the α -decay:



Now consider step 9, the decay of $^{214}_{82}\text{Pb}$, a radioactive isotope of lead, to $^{214}_{83}\text{Bi}$. On this occasion the mass number of the parent, $^{214}_{82}\text{Pb}$, and that of the daughter $^{214}_{83}\text{Bi}$, are both 214, whereas the atomic number of the daughter is one *more* than that of the parent. Thus, in order for the equation to balance, we need another decay product with a zero mass number a charge of minus 1. The electron meets these criteria, thus the decay is the β^- -decay:



The decay processes at the other stages can be found in a similar way.

Table 3.5 The radioactive decay scheme for uranium-238.

Step	Parent	Decay	Daughter
1	$^{238}_{92}\text{U}$	α	$^{234}_{90}\text{Th}$
2	$^{234}_{90}\text{Th}$	β^-	$^{234}_{91}\text{Pa}$
3	$^{234}_{91}\text{Pa}$	β^-	$^{234}_{92}\text{U}$
4	$^{234}_{92}\text{U}$	α	$^{230}_{90}\text{Th}$
5	$^{230}_{90}\text{Th}$	α	$^{226}_{88}\text{Ra}$
6	$^{226}_{88}\text{Ra}$	α	$^{222}_{86}\text{Rn}$
7	$^{222}_{86}\text{Rn}$	α	$^{218}_{84}\text{Po}$
8	$^{218}_{84}\text{Po}$	α	$^{214}_{82}\text{Pb}$
9	$^{214}_{82}\text{Pb}$	β^-	$^{214}_{83}\text{Bi}$
10	$^{214}_{83}\text{Bi}$	β^-	$^{214}_{84}\text{Po}$
11	$^{214}_{84}\text{Po}$	α	$^{210}_{82}\text{Pb}$
12	$^{210}_{82}\text{Pb}$	β^-	$^{210}_{83}\text{Bi}$
13	$^{210}_{83}\text{Bi}$	β^-	$^{210}_{84}\text{Po}$
14	$^{210}_{84}\text{Po}$	α	$^{206}_{82}\text{Pb}$ (stable)

Task 2

Taking a value for c , the speed of light, of $3.00 \times 10^8 \text{ m s}^{-1}$ gives a corresponding energy of

$$E = mc^2 = 9.2 \times 10^{-29} \text{ kg} \times (3.00 \times 10^8 \text{ m s}^{-1})^2 = 8.3 \times 10^{-12} \text{ J}$$

Note that this is a very small amount of energy (compare it with the kinetic energy of a running child, which is about 1 J). However, this is only the energy corresponding to the decay of a single atom.

Activity 3.3 Tossing coins and throwing dice

Tasks 1 and 2

The completed graph is shown in Figure 3.22.

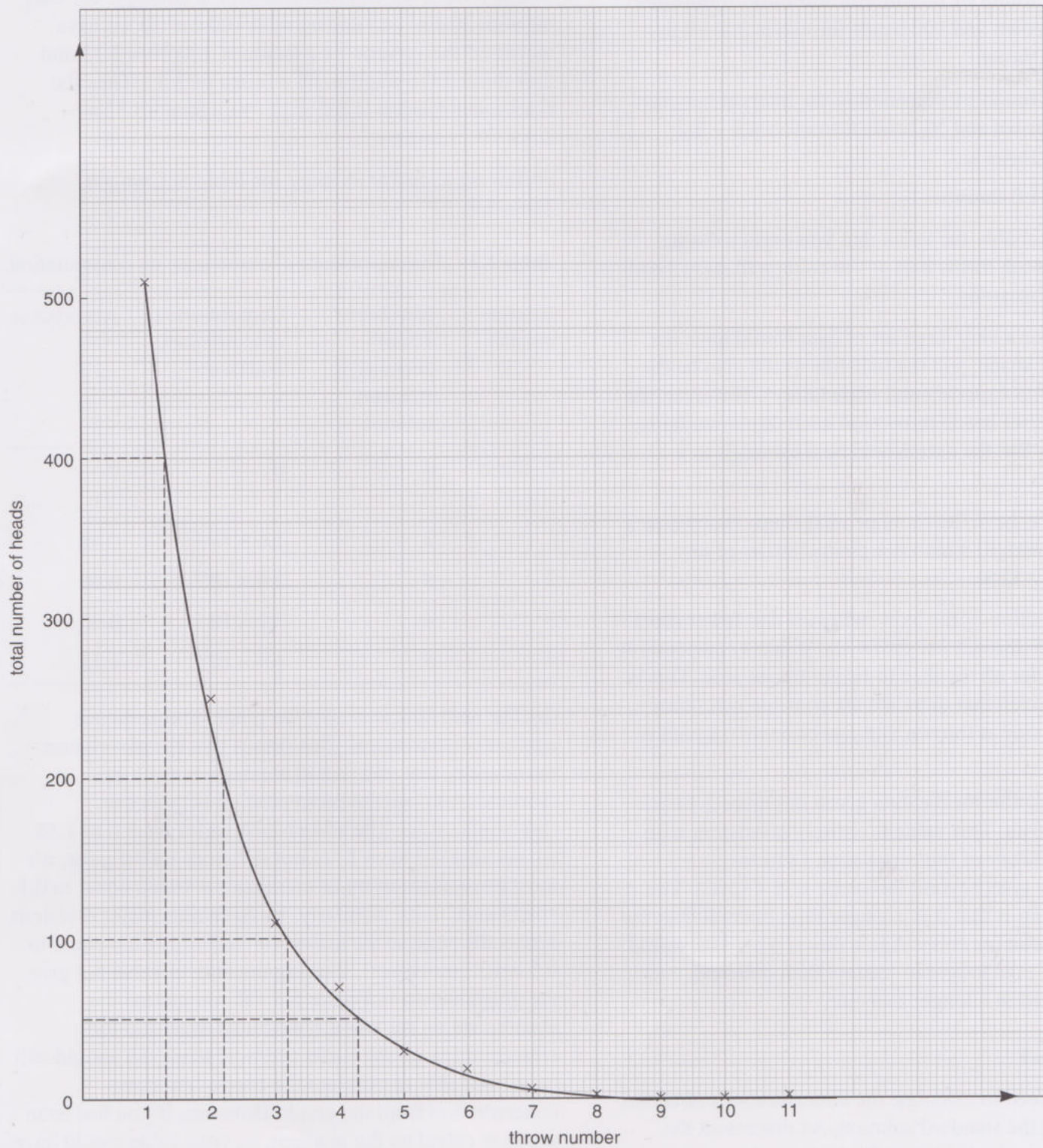


Figure 3.22 The total number of heads thrown at each throw.

Task 3

(a) The half-life found by considering the interval between 400 heads thrown and 200 heads thrown is $2.2 - 1.3$ throws = 0.9 throw.

The half-life found by considering the interval between 200 heads thrown and 100 heads thrown is $3.2 - 2.2$ throws = 1.0 throw.

The half-life found by considering the interval between 100 heads thrown and 50 heads thrown is $4.3 - 3.2$ throws = 1.1 throws.

The mean is 1.0 throws.

(b) This is probably the value that you expected – on average, for each throw half of the coins will show heads and thus be discarded.

(c) A graph for a very large set of data would have the same general shape, and the half-life would also be the same (to within experimental uncertainty). However, the actual number of heads thrown on each throw would be much greater and the data would be a closer fit to a smooth curve.

(d) The graphs have the same general shape (known as a falling exponential) which is a characteristic of any process of this kind.

Careful inspection of the axes indicates that fewer sixes than heads are thrown per throw and it takes considerably more throws for all the dice to be put to one side. This is reasonable, given that there are six possible outcomes every time a dice is thrown, but just two when a coin is tossed.

The mean half-life for the data given in Figure 3.14 is about 3.9 throws. This value is entirely consistent with the detailed theory of this simulation (which unfortunately goes beyond the scope of SXR103). The important thing is that it is a longer time interval than the equivalent value for coins. Again this is what you would expect – dice take longer to ‘decay’ than coins do, and this is reflected in a longer half-life.

Activity 5.1

(a) First we need to calculate the concentrations obtained after diluting the standard solution. As discussed, the simplest way to do this type of calculation is to determine by what factor the solution has been diluted. In solution 1 we start off with 20 cm^3 ($2 \times 10 \text{ cm}^3$) and end up with 100 cm^3 so we have increased the volume by a factor of $100/20 = 5$ and thus diluted the solution by a factor of 5.

Thus the final concentration is $8.0 \times 10^{-5} \text{ mol litre}^{-1}/5 = 1.6 \times 10^{-5} \text{ mol litre}^{-1}$.

Whilst this method of calculation works well for volumes that divide easily into 100 cm^3 , what about volumes like 30 cm^3 ? Well, we can follow a similar strategy. We start off with 30 cm^3 and end up with 100 cm^3 so we have increased the volume by a factor of $100/30 = 3.33$ and thus diluted the solution by a factor of 3.33. Thus the final concentration is $8.0 \times 10^{-5} \text{ mol litre}^{-1}/3.33 = 2.4 \times 10^{-5} \text{ mol litre}^{-1}$.

Following a similar strategy for Table 5.2, we get Table 5.3.

Table 5.3 The dependence of absorbance on concentration.

Solution number	Number of 10 cm^3 portions of standard solution	Concentration of aluminium/ mol litre^{-1}	Absorbance
Distilled water	0	0	0
1	2	1.6×10^{-5}	0.40
2	3	2.4×10^{-5}	0.60
3	4	3.2×10^{-5}	0.90
4	5	4.0×10^{-5}	1.15

(b) The next task is to plot the graph of absorbance *against* concentration. This means that the absorbance, the quantity that you would measure, is the y axis (vertical axis) and the concentration is the x axis (horizontal axis). The absorbance varies from 0 to 1.15 so, because we have 13 major divisions up the page, it's most convenient to make one major division equal to 0.1 absorbance units. Similarly, we have nine major divisions across the page so we make one major division equal to $0.5 \times 10^{-5} \text{ mol litre}^{-1}$. Plotting the data in Table 5.3 give the graph shown in Figure 5.3. (If you plotted the absorbance on the x axis (horizontal axis) and the concentration on the y axis (vertical axis) you should still have been able to measure the correct unknown concentration from the graph. However, if you had been asked to calculate the gradient, m , your value would have been $1/m$.)

The graph shows that the points all lie in a straight line, within experimental error, thus justifying our request for you to draw a straight line. We force the line to go through the origin, because we know that if there is no

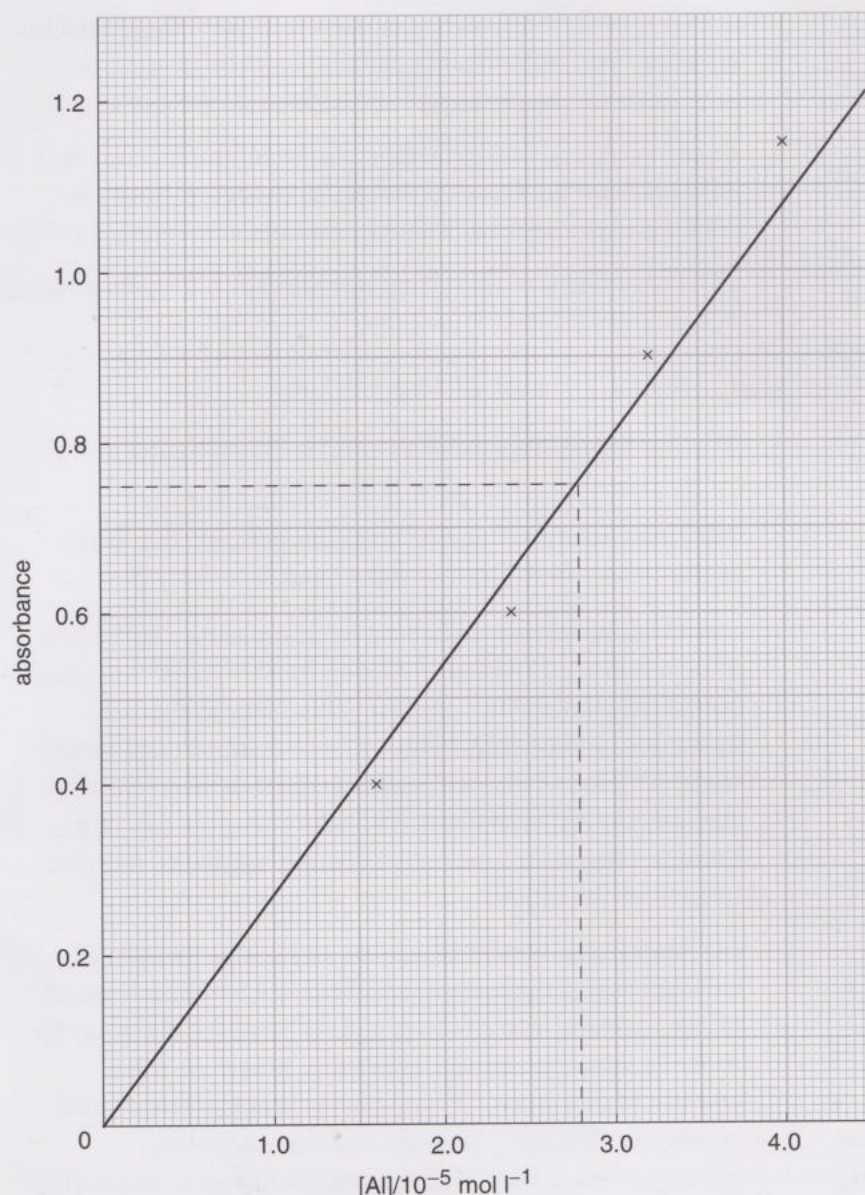


Figure 5.4 Plot of absorbance against concentration of aluminium.

aluminium present, the absorbance (arising from the aluminium) must be zero. The straight line confirms that the absorbance is *directly* proportional to the concentration, that is, we can write:

$$A = b[\text{Al}^{3+}]$$

where b is a constant and the gradient of your graph.

(c) The next step is to find the concentration of aluminium in an unknown water sample that has an absorbance of 0.75. Figure 5.4 shows that the point on the line that has an absorbance of 0.75 has a concentration of $2.8 \times 10^{-5} \text{ mol litre}^{-1}$. Depending upon where you draw your line the value you obtain could be anything between 2.7 and $2.9 \times 10^{-5} \text{ mol litre}^{-1}$.

The final stage is to convert the concentration of aluminium from mol litre^{-1} into p.p.m. A concentration of $2.8 \times 10^{-5} \text{ mol litre}^{-1}$ tells us that there is $2.8 \times 10^{-5} \text{ mol}$ of aluminium in 1 litre of water. The relative atomic mass

of aluminium is 27.0 thus 1 litre contains $27.0 \times 2.8 \times 10^{-5} = 7.6 \times 10^{-4} \text{ g}$ of aluminium (anything between 7.3 and $7.9 \times 10^{-4} \text{ g}$ is acceptable). Task A4 of the Activity B workbook tell us to assume that 1 litre has a mass of 1000 g. Thus $7.6 \times 10^{-4} \text{ g}$ of aluminium is contained in 1000 g. Put another way, $7.6 \times 10^{-1} \text{ g}$ of aluminium is contained in 1 000 000 g. Thus this corresponds to 0.76 p.p.m (0.73–0.79 p.p.m.).

- Have a look at the maximum acceptable concentrations of metals in drinking water, given in Table 1 on page 1 of the Activity B workbook. Is the unknown water sample safe to drink?
- The value you determined is greater than the maximum acceptable concentration of aluminium (0.2 p.p.m.) and so the sample is not fit to drink.

You are now ready to start the laboratory activities associated with this material.

Activity 6.1

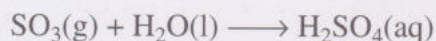
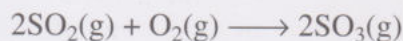
(a) (i) Two errors (or misprints) in the chemistry related to sulfate and hydrogen ions are:

- 'Sulfate ion (SO_4)' (page 1 column 3 line 18) should read 'Sulfate ion (SO_4^{2-})';
- 'Combines with hydrogen' (page 1 column 3 line 20) should read 'Combines with hydrogen ions, $\text{H}^+(\text{aq})$ '.

You may find other errors, for example the last paragraph in the third column on p. 4 states that sulfate ions are formed by oxidation of sulfur dioxide using ammonia, whereas the first column on p. 3 more correctly defines ammonia as the catalyst.

(ii) The conversion of sulfur dioxide (SO_2) into sulfate ion (SO_4^{2-}) involves the addition of O atoms, thus this is the definition of oxidation being used here.

(iii) The balanced equations are:



(b) (i) There are two processes — a wet phase and a gas phase.

The wet phase process is relatively fast and involves oxidation of $\text{SO}_2(\text{g})$ in clouds followed by reaction with water to make sulfuric acid (see (a) (iii) above). The acid is then deposited in rainfall.

The 'gas phase' process is much slower, i.e. $\text{SO}_2(\text{g})$ can be carried hundreds of kilometres unchanged and slowly reaches the ground unconverted.

(ii) The main source of sulfur dioxide is power stations. Gases emitted from power stations are treated with limestone (CaCO_3), which removes sulfur dioxide, forming calcium sulfate (CaSO_4).

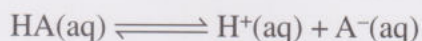
(iii) Two other gases that give rise to acid rain are nitric oxide (NO) and nitrogen dioxide (NO_2). These are given off by power stations and engine exhausts.

(iv) In Section 5.2 you were told that if the hydrogen ion concentration is written as 1.0×10^{-n} mol litre $^{-1}$, then the pH is simply equal to n . Thus, if the rain falling over the Pennines has a pH of 3.5, the hydrogen ion concentration of this rain is $1.0 \times 10^{-3.5}$ mol litre $^{-1}$. If the water in the clouds is ten times more acidic than this, then the hydrogen ion concentration in the clouds is $10 \times 1.0 \times 10^{-3.5}$ mol litre $^{-1}$. This can be written as $1.0 \times 10^{1-3.5}$ mol litre $^{-1} = 1.0 \times 10^{-2.5}$ mol litre $^{-1}$. So the pH of water in the clouds is 2.5.

(c) (i) A study of diatoms in lakes indicated that lakes in Scotland and Scandinavia did not become acidic until well after the beginning of the Industrial Revolution.

Other evidence is the disappearance of brown trout and salmon from lakes and rivers in Scandinavia since the 1950s and the decline of European forests in the mid 1970s.

(ii) These terms were defined in Section 5.2. A weak acid, HA, is only slightly dissociated:



That is, the equilibrium lies well to the left-hand side.

A strong acid HA is nearly fully dissociated, so effectively only H^+ and A^- ions are present, that is the equilibrium lies far to the right hand.

(d) (i) Most fish are poisoned by aluminium (Al^{3+} ions). Soil in lake and river beds is rich in Al^{3+} but it is strongly chemically bound to complex molecules in the soil. However, sulfate ions (SO_4^{2-}) from acid rain can detach these Al^{3+} ions from soil and transport them to water. Apparently Al^{3+} ions interfere with fish gills causing them to clog with mucus, preventing uptake of oxygen needed in the blood.

(ii) Acid (rain) waters leach out soil nutrients such as calcium and magnesium and liberate Al^{3+} ions. Lack of nutrients leads to starvation and Al^{3+} can be taken up by roots causing poisoning. There are, however, other complicating issues. $\text{SO}_2(\text{g})$ itself damages leaves and needles by blocking stomata and thus inhibiting photosynthesis. Ozone (from engine exhausts indirectly) may, in conjunction with $\text{SO}_2(\text{g})$, cause accelerated shoot growth, thus disrupting the growth balance. Ozone and SO_2 may also reduce the frost resistance of trees. These, and other pollutants appear to encourage harmful fungi and pests.

Activity 6.2

A typical summary, in 150 words is given below. Don't worry if your summary is different: it often depends upon what you think are the key points.

'Acid rain has become a global problem. The main culprit is sulfur dioxide from burning fossil fuel. This is oxidized then reacts with water to give sulfuric acid. Nitric oxide and nitrogen dioxide given off by power stations and car exhausts also forms nitric acid. The cause of acidification of lakes was confirmed by examining diatoms in their sediments, which are sensitive to acidity. The lake only became acidic after the Industrial Revolution. Fish

depletion in lakes and rivers is caused by the acid rain releasing aluminium from the soil into the rivers, resulting in clogging of fish gills. Tree depletion is caused by acidic waters washing out nutrients from soils and increasing the uptake of aluminium by the roots. Strict control of emissions in the USA and Europe has reduced the problem, as has treating the waste gases with limestone. Nevertheless recovery will take a long time.'

Activity 6.3

The relationship between atmospheric pollutants and damage to plants and animals is discussed below (in 396 words). Notice that an introduction has been given to contextualize some of the pollutants. The use of headings enables you to gather up all the information relating to one pollutant.

Introduction

Acid rain leads to groundwater with a pH in the region of 4. This leads to the release of aluminium from the soil. The main cause of acid rain is sulfur dioxide.

Many thin soils in Europe and North America have been acidic since they were formed, and this was first thought to be the cause of acid lakes. However, in 1980 a study on diatoms, which are very sensitive to acidity, suggested that the lake only became acidic after the Industrial Revolution. This is further supported by the fact that fish (brown trout and salmon) have disappeared from many Scandinavian lakes since the 1950s. Salmon in Southern Norway have decreased since 1970, and a similar pattern is observed in Scotland and Canada.

Pollutants and their action

Aluminium

All soils contain aluminium ions, which are locked up. Acidification releases aluminium ions to rivers and this interferes with fish gills, causing them to clog with mucus. This reduces the fishes' ability to take up oxygen, and they eventually die. Aluminium and hydrogen ions also effect the ecology of lakes which become covered with green algae and moss that remove nutrients, and the number of species in the lake consequently decreases.

In the Alps, fir trees started to lose needles in the mid 1970s. In 1986 29% of trees in the Netherlands were moderately defoliated or worse, 20% in West Germany, 16% in Czechoslovakia and Switzerland, and 29% in Britain. Acidic water leaches out nutrients (magnesium and calcium) and liberates aluminium. The trees are starved/poisoned when aluminium is taken up by their roots.

Sulfur dioxide

This can directly damage leaves and needles on trees by blocking stomata. Britain has a long history of pollution from coal burning. In 1952 a smog hung over London for a week, killing 4000 people. Acidic particles (pH 1.6) irritate bronchial tubes, leading to choking and heart attacks.

Ozone

This is a good thing in the stratosphere but a danger near the ground because it triggers asthma and leads to smog. It is also poisonous to trees, especially in conjunction with sulfur dioxide and other pollutants that reduce the frost hardiness of trees and encourage fungi and pests. It is formed in vehicle exhausts.

Ammonia

This can react with sulfur dioxide to give ammonium sulfate which coats the surface of vegetation. When washed out with rain, ammonium sulfate forms sulfuric acid and nitric acid in the soil. Too much nitrogen stresses plants through too much growth.

Answers to self-assessment questions

Question 2.1

Magma, such as volcanic lavas, that erupt at the Earth's surface will cool quickly so the crystals that form will be small, forming a fine-grained igneous rock.

On the other hand, magmas that are intruded deep in the Earth's crust will cool slowly giving large crystals time to form.

Question 2.2 Statement C is false whereas all the others are true.

Question 2.3

(a) High, because a large amount of energy is needed to transport the pebbles and boulders to the place where they are deposited.

(b) Low, because very little energy is needed to move fine clay particles.

(c) Medium, because sand grains are of a medium size.

Question 2.4

(a) Rocks formed by regional metamorphism possess a foliation due to the alignment of platy minerals that grew during compression. Rocks formed by contact metamorphism do not show foliation.

(b) Contact metamorphic rocks are confined to a narrow band in the rocks surrounding an igneous intrusion. Regional metamorphism occurs in vast tracts of land as a result of mountain building.

Question 3.1 The typical sizes of a nucleus and of an atom are 10^{-14} m and 10^{-10} m respectively. So,

$$\frac{\text{diameter of nucleus}}{\text{diameter of atom}} = \frac{10^{-14} \text{ m}}{10^{-10} \text{ m}}$$

Both diameters are measured in metres so the units cancel out.

$$\frac{\text{diameter of nucleus}}{\text{diameter of atom}} = \frac{10^{-14}}{10^{-10}}$$

This fraction can be simplified, because $10^{-14}/10^{-10} = 10^{-4}$.

The question asks for the answer to be expressed as $1/(\text{a number})$, and this is found using the fact that $10^{-n} = 1/10^n$. So $10^{-4} = 1/10^4$ and we can write

$$\frac{\text{diameter of nucleus}}{\text{diameter of atom}} = \frac{1}{10^4} = \frac{1}{10\,000}$$

Question 3.2 We start by rearranging the diffraction equation (Equation 3.4) to make λ the subject,

$$\lambda = \frac{d \sin \theta_n}{n}$$

(See *SGSG* page 370 for more on how to rearrange equations.)

In this case, only the first order of diffraction is of interest, hence $n = 1$:

$$\lambda = d \sin \theta_n$$

The wavelengths can be found using this equation.

The wavelength of light that produces first order diffraction at 19.8° is

$$\begin{aligned} \lambda &= 1.667 \times 10^{-6} \text{ m} \times \sin 19.8^\circ \\ &= 1.667 \times 10^{-6} \text{ m} \times 0.339 \\ &= 5.65 \times 10^{-7} \text{ m} \end{aligned}$$

to three significant figures. (See *SGSG* pages 349–352 for more on significant figures.)

The wavelength of light that produces first order diffraction at 23.4° is

$$\begin{aligned} \lambda &= 1.667 \times 10^{-6} \text{ m} \times \sin 23.4^\circ \\ &= 1.667 \times 10^{-6} \text{ m} \times 0.397 \\ &= 6.62 \times 10^{-7} \text{ m} \end{aligned}$$

to three significant figures.

Question 3.3

$\frac{\text{dose equivalent from Hiroshima and Nagasaki}}{\text{dose equivalent from 'Rocks and radioactivity'}}$

$$= \frac{5 \text{ Sv}}{2 \mu\text{Sv}} = \frac{5 \text{ Sv}}{2 \times 10^{-6} \text{ Sv}} = 2.5 \times 10^6$$

Question 3.4

$$\frac{m_p}{m_e} = \frac{1.673 \times 10^{-27} \text{ kg}}{9.11 \times 10^{-31} \text{ kg}} = 1840 \text{ to three significant figures.}$$

$$\text{So } m_e = \frac{1}{1840} m_p.$$

Question 3.5

(a) The electrons in the atom can only have specific values of energy (i.e. specific energy levels). When substance X is heated, the electrons gain energy so are excited to higher energy levels. At these higher energy levels, the electrons are less stable than at lower energy levels, so the electrons immediately tend to lose energy and fall back to lower levels. They give up the excess energy in the form of photons of electromagnetic radiation, detected as spectral lines. The photon energy corresponding to each spectral line is the same as the *difference* in energy between the two energy levels involved in the transition. Photon energy and wavelength are related by the equation

$$E_{\text{ph}} = \frac{hc}{\lambda}$$

so each different transition between energy levels gives rise to a characteristic wavelength. (See Section 3.3.)

(b) We have $E_{\text{ph}} = (E_3 - E_2) = 3.03 \times 10^{-19} \text{ J}$

$$\text{and } E_{\text{ph}} = \frac{hc}{\lambda}$$

(Equation 3.2). Rearranging the latter gives

$$\lambda = \frac{hc}{E_{\text{ph}}}$$

So

$$\lambda = \frac{6.63 \times 10^{-34} \text{ J s} \times 3.00 \times 10^8 \text{ m s}^{-1}}{3.03 \times 10^{-19} \text{ J}} = 6.56 \times 10^{-7} \text{ m}$$

to three significant figures. Converting into nanometres gives $\lambda = 656 \text{ nm}$.

Question 3.6 We have $\sin \theta_n = \frac{n\lambda}{d}$ (Equation 3.4). Rearranging gives

$$d = \frac{n\lambda}{\sin \theta_n}$$

So

$$d = \frac{1 \times 633 \times 10^{-9} \text{ m}}{\sin(30.4^\circ)} = 1.25 \times 10^{-6} \text{ m}$$

to three significant figures.

Question 3.7 The acceleration is given by the gradient of the graph in Figure 3.23:

$$\text{gradient} = \frac{\text{rise}}{\text{run}} = \frac{(30 - 5) \text{ m s}^{-1}}{(3.0 - 0.5) \text{ s}} = \frac{25 \text{ m s}^{-1}}{2.5 \text{ s}} = 10 \text{ m s}^{-2}$$

Thus the stone's average acceleration is 10 m s^{-2} .

(See *SGSG* p. 385 for how to calculate the gradient, and Section 3.5.7 for a discussion of uncertainties.)

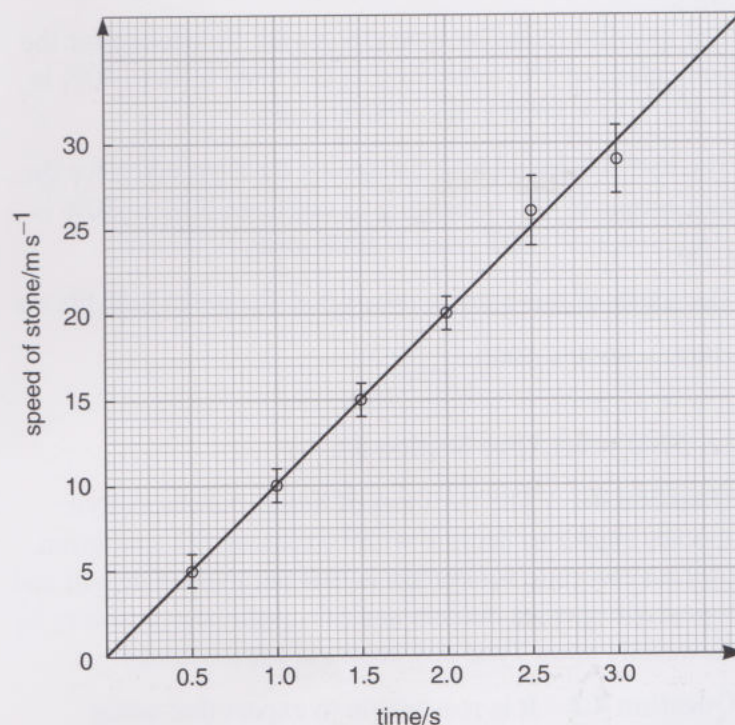


Figure 3.23 A completed plot of the data in Table 3.4.

Question 3.8 Radioactive decay is a random process, and a radioactive count of value n has an associated uncertainty of $\sqrt{n}$ (see Section 3.5.7). The uncertainties in Table 3.6 have been rounded to the nearest whole number above (i.e. the whole number *greater* than the calculated value).

Table 3.6 Completed Table 3.5.

Elapsed time/min	Counts per minute	Uncertainty in count rate per minute
1	403	21
2	320	18
3	249	16
4	201	15
5	157	13
6	128	12
7	109	11
8	89	10
9	71	9
10	66	9

Question 3.9 The half-life found by considering the time taken for the percentage of ^{14}C remaining to fall from 100% to 50% is $(5.8 - 0) \times 1000 \text{ years} = 5800 \text{ years}$.

The half-life found by considering the time taken for the percentage of ^{14}C remaining to fall from 80% to 40% is $(7.8 - 1.8) \times 1000 \text{ years} = 6000 \text{ years}$.

The half-life found by considering the time taken for the percentage of ^{14}C remaining to fall from 60% to 30% is $(10.2 - 4.3) \times 1000 \text{ years} = 5900 \text{ years}$.

The mean of these three values (i.e. the mean half-life) is

$$\frac{5800 + 6000 + 5900}{3} \text{ years,}$$

i.e. 5900 years. (See Section 3.5.5.)

Question 4.1 The numbers of lichen and bryophyte species appear to decline with increasing SO_2 pollution. (All we can state for certain is that numbers of lichen and bryophyte species decline and SO_2 pollution increases as Newcastle upon Tyne city centre is approached.)

Question 4.2 It is reasonable to expect that water pollution will kill other animals and plants.

Question 4.3

(a) At pH 7.5–8.5, molybdenum and sulfur are most available, whereas iron and phosphorus are least available. (Fair amounts of magnesium, calcium and nitrogen are also available, but there is comparatively little boron, copper, zinc or manganese.)

(b) In very acid soils of pH 3.5–4.5, manganese and iron are most available, but hardly any of the other minerals listed are available in significant amounts.

Question 4.4 Channelled wrack does not occur on very exposed shores. Limpets occur on both sheltered and exposed shores, where they show roughly the same upper limit of distribution. However, limpets extend farther down the shore in exposed habitats.

Question 4.5 As exposed surfaces and organisms lose water by evaporation, the concentration of salt in the remaining water increases.

Question 4.6

(a) There are no understorey or ground layers.

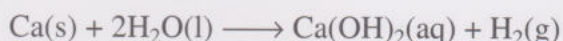
(b) One possibility is that the tree canopy is so dense throughout the year in the spruce plantation that hardly any light penetrates to the ground layer, whereas there are times of the year when light can penetrate through the tree canopy in deciduous woodland.

Question 4.7

(a) $10 \text{ m} \times 10 \text{ m}$ quadrats would probably be appropriate for recording tree species and numbers in a woodland. The moss layer in a lawn could be recorded using quadrats as small as $10 \text{ cm} \times 10 \text{ cm}$ (mosses are *much* smaller than flowering plants and grasses).

(b) Greenfly spend a lot of their life with their mouth parts stuck inside plants, but they are mobile for periods of time. When attached they can be counted per unit area, but when they are mobile, traps and timed samples are appropriate.

Question 5.1



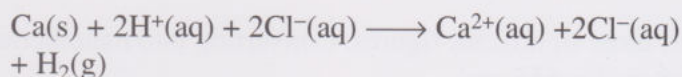
Question 5.2 At the atomic/molecular level, the equation may be interpreted as one atom of magnesium reacts with two hydrogen ions contained in an aqueous solution to form a magnesium ion and a molecule of hydrogen gas. However, a more useful interpretation is that one mole of solid magnesium atoms reacts with two moles of aqueous hydrogen ions to form one mole of aqueous magnesium ions with the evolution of one mole of gaseous hydrogen molecules.

Question 5.3

(a) The simple balanced equation is:



(b) The equation can also be written to highlight the ionic nature of this reaction:



The equation may be written without the chloride ions because these appear on both sides of the equation, i.e. they are spectator ions:



Question 5.4

(a) No reaction was observed when lead was dipped into either tin nitrate or zinc nitrate solutions, so lead is the least reactive of the three metals. Zinc, on the other hand, displaces both tin from tin nitrate solution and lead from lead nitrate, indicating it is the most reactive of the three metals. The position of tin as less reactive than zinc and more reactive than lead is confirmed by the results — that it displaces lead from lead nitrate solution, but does not react with zinc nitrate. The order of reactivity is therefore $\text{Zn} > \text{Sn} > \text{Pb}$.

(b) The balanced equation for the reaction of lead nitrate solution with tin is



Note that the nitrate ion, $\text{NO}_3^{-}(\text{aq})$, is a spectator ion and does not participate in the reaction, so may be left out of the ionic equation.

Question 5.5 The heat required q may be calculated from the equation (given in Section 4 of the Activity E workbook) $q = cm\Delta T$ and using the value for c given in the workbook:

$$q = 4.2 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1} \times 50 \text{ g} \times 20 \text{ }^{\circ}\text{C} = 4.2 \times 10^3 \text{ J} = 4.2 \text{ kJ}$$

Appendix 1

Sines, cosines and tangents

Look at the triangle drawn in Figure A1.1a. One angle of this triangle is a right angle and is therefore equal to 90° . The side of the triangle facing this is the longest of the three sides and is known as the hypotenuse^G. For convenience, it is labelled *hyp* in the diagram. Another of the angles of the triangle is labelled θ . The side of the triangle that is opposite the angle θ is labelled accordingly as *opp*, and the side adjacent to the angle is labelled *adj*.

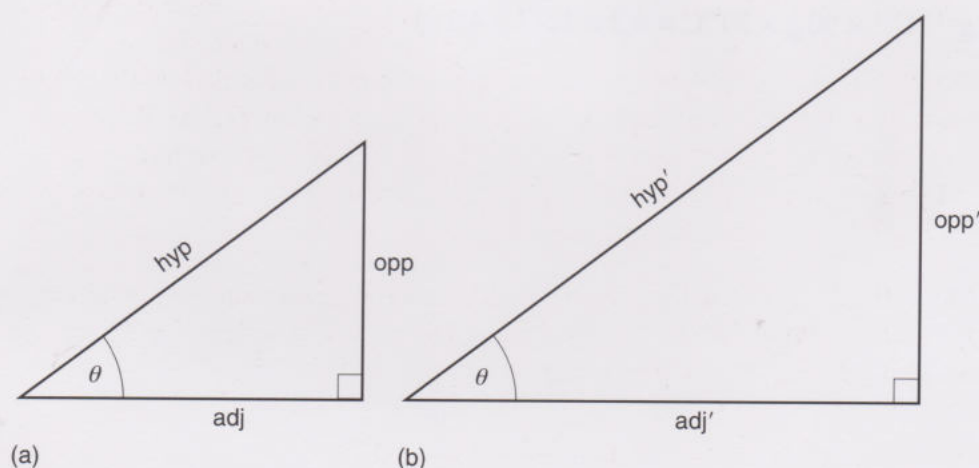


Figure A1.1 (a) and (b) show two triangles of different sizes, but with the same angles as each other.

A similar triangle can be drawn, with the same angles as the first triangle, but with the length of each side increased, as shown in Figure A1.1b. The sides of this triangle are labelled with the same abbreviations as the first, but with the symbol ' (prime) added.

Question A1.1 Each of the angles of the triangle shown in Figure A1.1a is the same as the corresponding angle of the triangle in Figure A1.1b. Measure the lengths of the sides of these triangles (*hyp*, *opp*, *adj* in Figure A1.1a and *hyp'*, *opp'*, *adj'* in Figure A1.1b). What are the values of

$$\left(\frac{\text{opp}}{\text{hyp}}\right), \left(\frac{\text{adj}}{\text{hyp}}\right) \text{ and } \left(\frac{\text{opp}}{\text{adj}}\right)$$

for the triangle in Figure A1.1a? How do these compare with the values

$$\left(\frac{\text{opp}'}{\text{hyp}'}\right), \left(\frac{\text{adj}'}{\text{hyp}'}\right) \text{ and } \left(\frac{\text{opp}'}{\text{adj}'}\right)$$

for the triangle in Figure A1.1b?

As you saw in Question A1.1, the relative lengths of the sides of a right-angled triangle depend on the angles in that triangle. Turning this argument around, the sizes of the angles in a right-angled triangle depend on the relative lengths of the sides of these triangles. This is such a useful fact that the values you have calculated above

are given special names: sine^G, cosine and tangent, often abbreviated to sin, cos and tan. For the angle θ shown in Figure A1.1:

$$\sin \theta = \left(\frac{\text{opp}}{\text{hyp}} \right)$$

$$\cos \theta = \left(\frac{\text{adj}}{\text{hyp}} \right)$$

$$\tan \theta = \left(\frac{\text{opp}}{\text{adj}} \right)$$

These three trigonometric functions, as they are known, are stored on your calculator. So, for instance, if you key into your calculator: 3 0 SIN or SIN 3 0, depending on the operation of the calculator, it will display a value of 0.5. Therefore the sine of 30° is 0.5. In any right-angled triangle, in which the length of one side divided by the length of the hypotenuse is 0.5, the angle opposite to the side in question will be 30° . Sine, cosine and tangent functions refer to relative lengths of the sides of a right-angled triangle, and so are independent of the actual size of the triangle in question, as you saw in Question A1.1, and they have no units.

Answer to Question A1.1 The lengths of the side in the triangles shown are opp = 3.0 cm, adj = 4.0 cm, hyp = 5.0 cm, opp' = 4.5 cm, adj' = 6.0 cm and hyp' = 7.5 cm. The values calculated for the smaller triangle are

$$\frac{\text{opp}}{\text{hyp}} = \frac{3.0}{5.0} = 0.6, \quad \frac{\text{adj}}{\text{hyp}} = \frac{4.0}{5.0} = 0.8 \quad \text{and} \quad \frac{\text{opp}}{\text{adj}} = \frac{3.0}{4.0} = 0.75.$$

The values for the larger triangle are

$$\left(\frac{\text{opp}'}{\text{hyp}'} \right) = \frac{4.5}{7.5} = 0.6, \quad \left(\frac{\text{adj}'}{\text{hyp}'} \right) = \frac{6.0}{7.5} = 0.8 \quad \text{and} \quad \left(\frac{\text{opp}'}{\text{adj}'} \right) = \frac{4.5}{6.0} = 0.75.$$

The relative lengths of the sides are therefore the same in triangles whose angles are the same.

In fact, the relative lengths of the sides of a triangle are the same for *any* triangle that has the same angles, however large or small the triangle is. But if we change the shape of the triangle, so that the angles are different, then the relative lengths of the sides of the triangle will be different too.

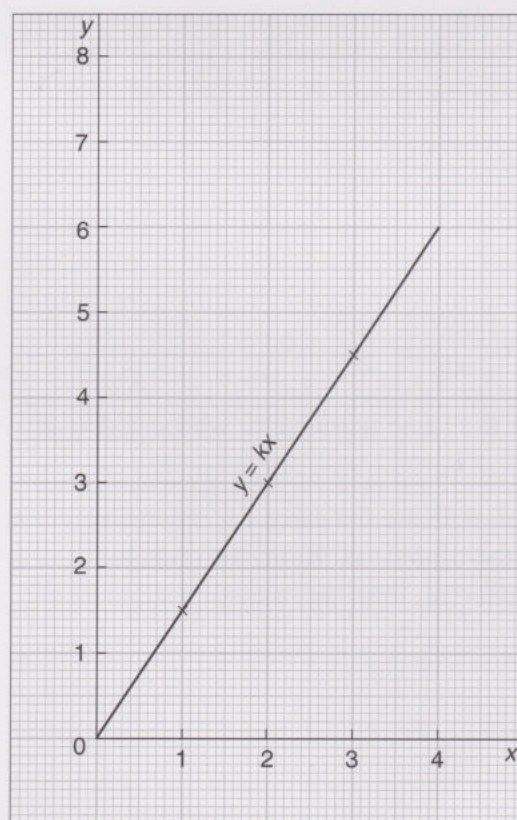
Appendix 2

The equation of a straight line

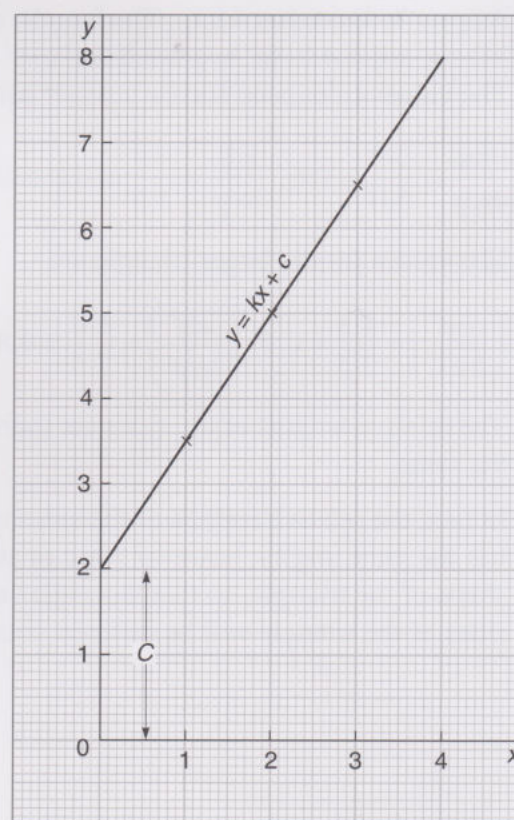
Imagine you are walking along a road at a constant speed. The distance travelled from your starting point will be proportional to the time taken, so in twice the time you will travel twice the distance. If we plotted a graph of distance travelled against time taken, we would get a straight line passing through the origin of the graph. In fact, if values for the two quantities on either side of *any* proportionality relationship (such as $y \propto x$, i.e. $y = kx$, where k is a constant) are plotted against each other on a graph, then the points will lie along a straight line that passes through the origin, as shown in Figure A2.1. The constant k is equal to the gradient of the graph, and can be calculated as described on page 385 of *SGSG*. Any equation that can be written in the form

$$y = kx$$

(often given as $y = mx$) is referred to as 'an equation of a straight line'.



(a)



(b)

Figure A2.1 (a) A graph showing values of a quantity y plotted against corresponding values of x and illustrating the proportionality $y \propto x$. The gradient of this graph is k , such that $y = kx$. (b) A graph showing values of a quantity y plotted against x and illustrating the equation $y = kx + c$. The gradient of this graph is k and it intercepts the vertical axis at $y = c$.

Now look at the line plotted on the graph in Figure A2.1b. This has the same gradient as the graph in Figure A2.1a, but the whole line has been moved vertically upwards by a distance c . So for any value of x , we can find the value of y by calculating kx as with the graph in Figure A2.1a, and then adding on an extra amount c , corresponding to the vertical shift. So the equation that allows us to calculate a value for y from a value of x is:

$$y = kx + c$$

This is the general form of the equation of a straight line^G. It is often given in a slightly different form, as $y = mx + c$, where m is the gradient.

From this equation you can see that, when $x = 0$, then $y = (k \times 0) + c$, and therefore $y = c$. So the value of c indicates the point at which the line intercepts the vertical axis. Such a graph still has a gradient of k (or m), however, as you can see by comparing the gradients of the two graphs in Figure A2.1. Figure A2.2 shows how a given straight line graph is related to its equation.

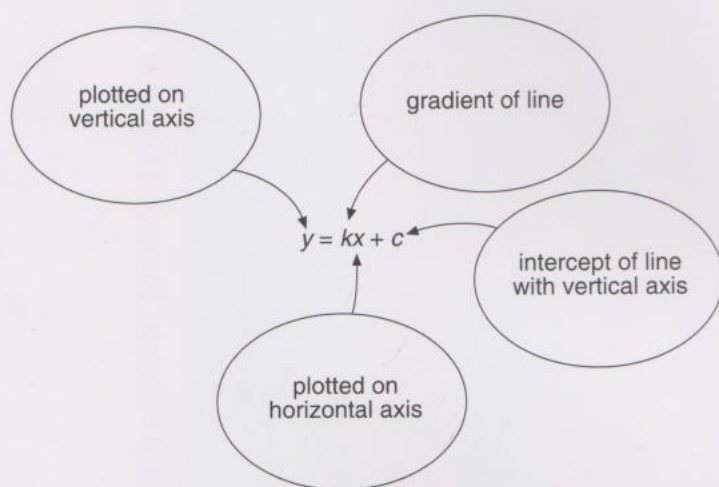


Figure A2.2 The equation of a straight line.

- If c was a negative number, how would the graph differ from that shown in Figure A2.1b?
- The graph would still be a straight line, and would slope in the same direction, but it would be shifted downwards so that it intercepted the vertical axis in the region where y is negative.

Appendix 3

Group		Group																							
I II		III	IV	V	VI	VII	0																		
Period 1								1	2																
								H	He																
Period 2		3	4																						
		Li	Be																						
Period 3		11	12																						
		Na	Mg																						
Period 4		19	20																						
		K	Ca																						
Period 5		37	38																						
		Rb	Sr																						
Period 6		55	56																						
		Cs	Ba																						
Period 7		87	88																						
		Fr	Ra																						

Table A3.1 Some relative atomic masses.

Element (symbol)	Relative atomic mass
Hydrogen (H)	1.01
Oxygen (O)	16.0
Aluminium (Al)	27.0
Sulfur (S)	32.1
Chlorine (Cl)	35.5
Copper (Cu)	63.5
Zinc (Zn)	65.4

Acknowledgements

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Figures

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INDEX

Entries and page numbers in **bold type** refer to terms defined in the glossary. Where the page number is given in *italics*, the indexed information is given mainly, or wholly, in a figure or table.

- absorbance 90, 91, 104–5
- acid rain 95–9, 106–7
- acids, reactivity 80–4
- air pollution 68, 69, 70
 - acid rain 95–9, 106–7
 - sulfur dioxide 69, 95, 106–7
- alkalis, reaction with hydrochloric acid 80–4
- alpha-decay (α -decay) 48**
- alpha-particle (α -particle) 48**
- aluminium 89–90, 94, 107
- anaphase 67**
- angle of diffraction 44, 45, 46, 57**
- animals
 - data collection 74, 78
 - fossilized 20, 22
 - see also* fauna; fish; invertebrates
- atomic mass 82, 117
- atomic nucleus 37, 47**
 - in radioactive decay 48–9, 50
- atomic number 47
- atoms 37
 - spectra 38–9, 41
- background radiation 57, 58
- beaches 34
- bed, sedimentary rocks 19
- beta-minus decay (β -decay) 48–9
- beta-particle (β -particle) 49
- biodiversity 68
- bryophytes (liverworts) 68, 69
- calcicoles, species 70, 71, 72**
- calcifuges, species 70, 71**
- calcite 8, 9, 17, 25**
- calibration 45**
 - colorimeter 89
 - diffraction grating 45–6, 101–2
- cell cycle 66
- cell division 66–7
- cell membrane 65, 67
- cell walls 65, 67
- cells 65
- centromere 66, 67
- chalk 17, 70
- chemical elements 37
 - isotopes 47
 - periodic table 116
 - spectra 42
- chemical equations 80–1, 83, 85, 89
- chemical equilibrium 81**
- chemical tests, unknown substances 88–9
- chloroplasts 65**
- chromatids 66, 67**
- chromosomes 66, 67**
- cleavage, in minerals 12–13, 14**
- coastline *see* sea-shore
- coin tossing 51–5, 103, 104
- collimator 44–5**
- colorimeter 89
- communication skills 93–4, 100
- concentration
 - absorbance effect 104–5
 - calculation 89–91
- continental collision zones *see* plate boundaries
- continuous spectrum 38**
- cosine 112–13
- cross-stratification 23–4**
- crystals, in igneous rocks 8, 10–12
- cytoplasm 65**
- data collection *see* quantitative data
- deposition, sedimentary process 17, 30–1, 33, 34
- dice throwing 51–5, 103, 104
- diffraction grating 43, 44**
 - calibration 45–6, 101–2
- dilution factors 89, 90, 104
- diploid 66**
- displacement reactions, metals 84–8**
- dissociation, acids 81**
- DNA 66, 67
- dose equivalent 58**
- Earth's crust, cross-section 7, 32
- earthworms 70, 71
- ecology 68–70**
 - see also* environment
- electrolysis 87
- electromagnetic radiation 42**
- electron antineutrino 49**
- electron structure 37**
- electrons 37**
 - energy levels 39–41
 - in oxidation 85
 - in radioactive decay 48–9
- elements *see* chemical elements
- emission spectrum 38, 39, 41, 45**
- endothermic reactions 87**
- energy levels**
 - in atoms 39–41
 - of photons 39–41, 42–3
- enthalpy change 87–8
- environment
 - analysis 88–91
 - damage to 68–70, 95–9
 - see also* habitats; pollution
- equation of a straight line 114–15
- equilibrium constant 81
- eras, geological 20, 21
- erosion 10, 17, 33
- error bars, on graphs 57
- eukaryotes 65
- excited states, energy levels 40–1
- exothermic reactions 87
- extraction, metals 86–7
- faults (in rocks) 30**
- fauna 68**
- feldspar 12, 15**
- fieldwork
 - data collection 77–8
 - ecology 70, 74–7
 - geology 27–31
 - sketches 27–8, 74, 75
- fish, water pollution 69, 107
- flora 68**
- foliation 25–6**
- formula unit 82**
- fossils 17, 20–2**
- fractures (in rocks) 30**
- gametes 66
- gamma-decay (γ -decay) 49, 58
- gamma-rays (γ -rays) 42, 59
 - biological effects 65–8
- Geiger counter 55, **57, 58–9, 62**
- geological time 20, 21
- glaciation 28, 34
- gneiss 26**
- grain size, rocks 26
- grassland 70
- grating spacing 44, 45**
- ground layer 74, 75**
- ground state, energy level 40, 41**
- habitats
 - preservation 68, 70
 - sea-shores 72–4
 - woodland 74–7
- half-life 50, 52, 59
- half-thickness, materials 59
- hand lens 9
- handheld spectroscope 45
- haploid 66
- heat
 - cause of metamorphism 24–5
 - effect on spectral lines 45
 - from chemical reactions 87–8
- helium atom 47, 48
- hydrochloric acid, reactivity 80–4
- hydrogen atom
 - ions 81–2
 - model 39–41
 - spectrum 60
- hypotenuse 112
- igneous rocks 10
 - classification 12–16
 - cooling rate 26
 - crystallization 10–12
 - metamorphism 24
- information skills 93–5
- interphase 66, 67**
- inverse square law 58**

invertebrates 70

see also earthworms; fish; limpets; wracks
ion 41
ionic compounds, reactions 83, 84–6, 89
ionization 41
isotopes 47
half-life 50, 52, 59
uranium 49–50

joints (in rocks) 30

journal articles 93, 94–9

lava flow 12, 13

lead, radiation absorber 59

lichens (mosses) 68, 69

limestone 9, 17, 25

soil pH 70, 72

limpets 73

lithospheric plates 7

lustre (in minerals) 12

magma 10–13, 24

mass number 47, 48

meiosis 66

metals

displacement reactions 84–8

emission spectra 45

extraction 86–7

gamma-ray absorption 59

hydrochloric acid reaction 80–4

oxidation 85

salts 84–6, 88–91

metamorphic rocks 10, 24–6

metamorphism 24–6

see also regional metamorphism

metaphase 67

mica 13, 14, 15, 25, 26

microscope 37

mineral nutrients, in soils 71, 72

minerals 8, 11, 15

cleavage 12–13, 14

identification 9, 12

mitosis 66–7

molar mass 82, 89, 90

mole, concentration 81, 82

monochromatic light 43, 45

mountains 7, 33

multicellular (organism) 65

mutation 66

nanometre 45, 46

neutrons 37, 47

nuclear decay *see* radioactive decay

nuclear membrane 65

nucleons 47

nucleus (of a cell) 65, 66, 67

nucleus (of an atom) *see* atomic nucleus

oceans 7

olivine 15

order of diffraction 44

ore extraction 86–7

organisms *see* animals; plants

oxidation reactions 83–4, 85

Periodic Table 84, 116

Periods, geological 20, 21

pH

of acids 82–3

of soil 70–2

photons 38, 39–41, 42–3, 49

photosynthesis 65

phytoplankton fossils 17

Planck constant 42

plants

alkaline-loving 70–2

cellular structure 65–7

data collection 74, 77–8

fossilized 20, 22

soil pH 71, 72

woodland 74–7

see also bryophytes; flora; lichens;

seaweed

plate boundaries 7, 24, 25, 33

point quadrat, fieldwork 77

pollen grain 66

pollution 68–70

acid rain 95–9, 106–7

see also air pollution; water pollution

precipitation reactions 88–9

principal quantum number 40

progeny cells 66, 67

prophase 67

protons 37, 47

in radioactive decay 48–9

pyroxene 12, 14, 15

quadrat, fieldwork 77

quantitative data

collection 77–8

in fieldwork 74

transect line 69, 75, 77

quantum 38

quartz 8, 15, 18, 24–5

radiation *see* background radiation;

electromagnetic radiation

radio waves 42

radioactive decay 48

counting 55–7, 62

safety 58–9

uranium 47, 48, 49, 50, 60, 102

raised beach 34

reactivity series 84, 85–6, 87

reduction reactions 83–4, 85, 86

regional metamorphism 24, 25, 33

relative atomic mass 82, 117

rivers, sedimentation 17, 33

rock cycle 32–3

rocks 7–8, 9–10

boulders 28, 29

see also igneous rocks; metamorphic

rocks; sedimentary rocks

salts, reactions 84–6, 88–91

sand 17, 18, 19, 23

sandstone 23, 24–5, 70

scanning tunnelling microscope 37

schist 26

scientific information 93–5

sea-level 33–4

sea-shore

fieldwork 70

habitats 72–4

rocks 19, 28–30

see also beaches

seaweed 73

sedimentary rocks 10, 16–17

composition 22–4

metamorphism 24–5

strata 19–20, 23, 28–31

sedimentation 17, 33

sievert, radiation dose 58

silicate minerals 8, 12, 13–14, 15

simulation, radioactive decay 51–5

sine 112–13

sketches, in fieldwork 27–8, 74, 75

slate 26

soil, pH 70–2

sorting (of sediments) 18

species *see* animals; organisms; plants

spectral lines 38, 39, 41, 42

wavelength measurement 43–4

spectrometer 44–5, 57

calibration 45–6

spectroscope 45

spectroscopy 42

spectrum 38

see also emission spectrum

standard solution 89–90

strata, in rocks 19–20, 23, 28–31

stratigraphic column 20, 21

stratigraphy 20

strong acid 81

sulfur dioxide 69, 95, 106–7

superposition 19

tangent 112–13

teamwork 94, 95

telescope 44, 45

telophase 67

texture, in rocks 9, 10–12

timescale, geological 20, 21

transect, data collection 69, 75, 77

tree layer 74, 75

trees 74–7, 107

uncertainty, decay counting 57, 60

unconformity 20, 31

understorey 74, 75

uranium 47, 48, 49, 50, 60, 102

volumetric flask 90

water pollution 68, 69, 70, 107

wave 42

wavelength 42, 43–4

weak acid 81

weathering 17, 18, 30, 33

woodland 74–7

wracks 73

zonation pattern 72, 73

zygote 66